Lecture Notes Package for

STEM Core ChemJam

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Includes Worksheets

Chemistry

Chemistry is the study of matter and how it interacts with other _____________ and/or _____________.

Matter and Energy

Matter is anything that has _______________ and occupies ______________.

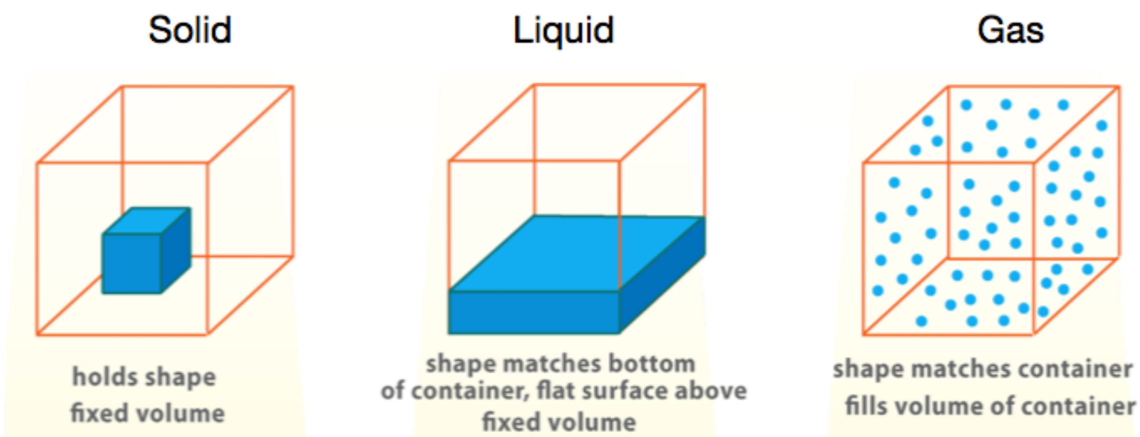
We can describe matter in terms of________________ __________________, those characteristics that can be determined without changing it into a different substance.

- Example: Sugar is white, tastes sweet, and can be crushed into powder. Crushing sugar does not change sugar into something else.

Matter can also be described in terms of its ______________ ***properties***. Chemical properties of substances describe *how they are converted to new substances* in processes called chemical reactions.

- Example: Caramelization of sugar

Matter is typically found in one of three different physical ___________ (sometimes called _________).



Example: Ice **Example: water** **Example: steam**

Changing the phase of matter, converting matter between solid, liquid, and gas is considered a *physical change* because the identity **does not** change.

- Examples of phase changes are: melting, boiling water to make steam, and melting an iron rod.

Energy

Energy is commonly defined as the ability to do ________.

Energy can be found in two forms, ______________ *energy* and _____________ *energy*.

Potential energy is ___________ energy; it has the potential to do work.

- An example of potential energy is water stored in a dam. If a valve is opened, the water will flow downhill and turn a paddle connected to a generator to create electricity.

Kinetic energy is the energy of _____________.

- Any time matter is moving, it has kinetic energy.

An important law that is central to understanding nature is: **matter will exist in the lowest possible energy state**. Another way to say this is “if matter can lose energy, it will always do so.”

Understanding Check: Kinetic Energy vs. Potential Energy

Which are mainly examples of ***potential energy*** and which are mainly examples of ***kinetic energy***?

a) A mountain climber sits at the top of a peak.
b) A mountain climber rappels down a cliff.
c) A hamburger sits on a plate.
d) A nurse inflates a blood pressure cuff.

Units of Measurements

Measurements consist of two parts – a ________________ and a____________________.

SI Units and Their Symbols

Quantity	SI Unit Name	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	Kelvin	K

Commonly Used Units and Their Symbols

Quantity	Unit Name	Symbol
Length	foot inch	ft in
Mass	gram pound	g lb
Volume	Liter	L
Temperature	Fahrenheit Celsius (or Centigrade)	^{0}F ^{0}C

Scientific Notation and Metric Prefixes

Scientific Notation

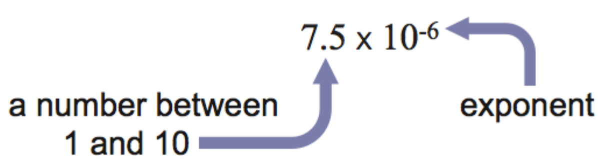
When making measurements, particularly in science and in the health sciences, there are many times when you must deal with very large or very small numbers.

Example: a typical red blood cell has a diameter of about 0.0000075 m.

In ________________ __________________(exponential notation) this diameter is written 7.5×10^{-6} m.

Values expressed in scientific notation are written as a number between _____ and _____ multiplied by a power of 10.

The superscripted number to the right of the ten is called an *exponent*.



- An exponent with a positive value tells you how many times to **multiply** a number by 10.

$$3.5 \times 10^4 = 3.5 \times 10 \times 10 \times 10 \times 10 = 35000$$

- An exponent with a negative value tells you how many times to **divide** a number by 10.

$$3.5 \times 10^{-4} = \frac{3.5}{10 \times 10 \times 10 \times 10} = 0.00035$$

Converting from Regular Notation to Scientific

1) Move the decimal point to the right of the first (right-most) non-zero number

- The exponent will be equal to the number of decimal places moved.

2) When you move the decimal point to the left, the exponent is positive.

$$35000 = 3.5 \times 10^{4}$$

$$285.2 = 2.852 \times 10^{2}$$

$$8300000 = 8.3 \times 10^{6}$$

3) When you move the decimal point to the right, the exponent is negative.

$$0.00035 = 3.5 \times 10^{-4}$$

$$0.0445 = 4.45 \times 10^{-2}$$

$$0.00000003554 = 3.554 \times 10^{-8}$$

Understanding Check: Convert each number into scientific notation.

a) 0.0144
b) 144
c) 36.32
d) 0.0000098

Converting from Scientific Notation to Regular Notation

You just learned how to convert from regular numerical notation to scientific notation. Now let's do the reverse; convert from scientific notation to regular notation.

Step 1: Note the value of the ***exponent***.

Step 2: The value of the exponent will tell you which direction ***and*** how many places to move the decimal point.

- If the value of the exponent is **positive**, remove the power of ten and move the decimal point that value of places to the ***right***.
- If the value of the exponent is **negative**, remove the power of ten and move the decimal point that value of places to the ***left***.

Example: Convert 3.7 x 10^5 into regular notation.

Step 1: Note the value of the ***exponent***: The exponent is **positive 5**.

Step 2: The value of the exponent will tell you which direction and how many places to move the decimal point.

- If the value of the exponent is **positive**, remove the power of ten and move the decimal point that value of places to the ***right***.
  - We will move the decimal point 5 places to the ***right***.

$$3.7 \longrightarrow 3.70000 \longrightarrow 370000$$

When the decimal point is **not shown** in a number, as in our answer, it is assumed to be ***after the right-most digit***.

Let's do another example: Convert 1.604 x 10^{-3} into regular notation.

Step 1: Note the value of the ***exponent***: The exponent is ***negative* 3**.

Step 2: The value of the exponent will tell you which direction and how many places to move the decimal point.

- If the value of the exponent is **negative**, remove the power of ten and move the decimal point that value of places to the ***left***.
    - We will move the decimal point 3 places to the ***left***.

$$1.604 \longrightarrow 001.604 \longrightarrow .001604$$

or 0.001604

Understanding Check: Convert the following numbers into regular notation.

a) 5.2789 x 10^{2}
b) 1.78538 x 10^{-3}
c) 2.34 x 10^{6}
d) 9.583 x 10^{-5}

Measurements and Significant Figures

There are three important factors to consider when making measurements:

1) **accuracy**
2) **precision**
3) **significant figures**

_____________________ is related to how close a measured value is to a true value.

Example: Suppose that a patient's temperature is taken twice and values of 98 °F and 102 °F are obtained. If the patient's true temperature is 103 °F, the second measurement is more *accurate*.

___________________ is a measure of reproducibility.

Example: Suppose that a patient's temperature is taken three times and values of 98 °F, 99 °F, and 97 °F are obtained. Another set of temperature measurements gives 90 °F, 100 °F, and 96 °F.

- The values in the first set of measurements are closer to one another, so they are more precise than the second set.

The quality of the equipment used to make a measurement is one factor in obtaining accurate and precise results. The ability of the human operator to correctly use the measuring device is another factor.

Significant Figures

One way to include information on the ______________ of a measured value (or a value that is calculated using measured values) is to report the value using the correct number of significant figures.

The precision of a measured value can be determined by the ___________ -most decimal place reported.

- The names and precision of the decimal places for the number **869.257** are shown below:

Digit in Number	8	6	9	2	5	7
Decimal Place Name	Hundreds (100's)	Tens (10's)	Ones (1's)	Tenths ($1/10^{th}$'s)	Hundredths ($1/100^{th}$'s)	Thousandths ($1/1000^{th}$'s)

A simple way to understand significant figures is to say that a digit is significant if we are _________ of its value.

Method for Counting Significant Figures

Measured and calculated values should be reported using significant figures. We can look at a numerical value and determine the number of significant figures as follows:

- If the decimal point is ____________, starting from the ***left***, count all numbers (including zeros) beginning with the first non zero number.
- If the decimal point is ____________, starting from the ***right***, count all numbers (including zeros) beginning with the first non zero number.
- When numbers are given in scientific notation, **do not** consider the power of 10, only the value before " **x 10^{n}**."

Example: If the botanist reported the age of the tree as **500 years**, how many significant figures are given?

Note that although the decimal point is implied to be after the right-most zero, it is **absent** (not shown explicitly), therefore we use the decimal point **absent** rule shown above; if the decimal point is **absent**, starting from the ***right***, count all numbers (including zeros) beginning with the first non zero number.

- We will start inspecting each digit from right (to left) as shown by the arrow.
- We will start counting when we get to the first non zero number.

500 ←

We do not count the first two zeros, but start counting at the **5**. Therefore, there is **one** significant figure present.

Example: If the botanist reported the age of the tree as **500.** years (note the decimal point present), how many significant figures are given?

Note that in this case, the decimal point is **present** (shown), therefore we use the decimal point **present** rule shown above; if the decimal point is **present**, starting from the ***left***, count all numbers (including zeros) beginning with the first non zero number. We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.

→ 500.

We begin with the **5**, then count all numbers ***including zeros***. In this case, the two zeros are also significant. Therefore there are **three** significant figures present.

Outside of the science fields, "**500**" and "**500.**" are generally thought of as equivalent, however, the use of significant figures tells us that when we write "**500.**" (with the decimal point present) we know that number one hundred times more precisely than when we write "**500**" (without the decimal point). We have precision to the "ones" decimal place in "**500.**" vs. precision to the "hundreds" place in "**500**".

Here are some other examples:

Example: How many significant figures are contained in 0.00045?

Note that in this case, the decimal point is **present** (shown). We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.

$$\longrightarrow 0.00045$$

We begin with the **4**, then count <u>all</u> numbers ***including zeros***. Therefore there are **two** significant figures present.

Example: How many significant figures are contained in 0.0002600?

Note that in this case, the decimal point is **present** (shown). We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.

$$\longrightarrow 0.0002600$$

We begin with the **2**, then count <u>all</u> numbers ***including zeros***. Therefore there are **four** significant figures present.

Example: How many significant figures are contained in 7080?

If the decimal point is **absent**, starting from the ***right***, count all numbers (including zeros) beginning with the first non zero number. We will start inspecting each digit from right (to left) as shown by the arrow. We will start counting when we get to the first non zero number.

$$7080 \longleftarrow$$

We do not count the first zero, but start counting at the **8,** and then count ***<u>all</u>*** *numbers **(including zeros)***. Therefore, there are **three** significant figures present.

Understanding Check: Specify the number of significant figures in each of the values below.

a) 23.5	f) 6200.
b) 0.0073000	g) 6200.0
c) 6.70	h) 0.6200
d) 48.50	i) 0.62
e) 6200	j) 930

Significant Figures in Scientific Notation

When numbers are given in scientific notation, **do not** consider the power of 10, only the value before "**x 10ⁿ**."

Examples: How many significant figures are contained in each of the values shown below?

a) $\boxed{5} \times 10^2$ **one** significant figure

b) $\boxed{5.0} \times 10^2$ **two** significant figures

c) $\boxed{5.00} \times 10^2$ **three** significant figures

When converting back and forth from standard numerical notation to scientific notation, the number of significant figures used **should not change**.

Understanding Check: Write each measured value in ***scientific notation***, being sure to use *the correct number of significant figures*.

a) 5047
b) 87629.0
c) 0.00008
d) 0.07460

Calculations Involving Significant Figures

When doing ________________ ____ _________________with measured values, the answer should have *the same number of significant figures* as the measured value with the least number of significant figures.

When doing ________________ ____ _________________with measured values, the answer should have the same *precision* as the least precise measurement (value) used in the calculation.

Example for Multiplication or Division:

- When doing ***multiplication or division*** with measured values, the answer should have ***the same number of significant figures*** as the measured value with the least number of significant figures.
- **Example:** If an object has a mass of 5.324 grams and a volume of 7.9 ml, what is its density?

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{\overbrace{5.324\text{ g}}^{\text{4 sig figs}}}{\underbrace{7.9\text{ ml}}_{\text{2 sig figs}}} = \underbrace{0.67\text{ g/ml}}_{\text{2 sig figs}}$$

Example for Addition or Subtraction:

- When doing ***addition or subtraction*** with measured values, the answer should have the same ***precision*** as the least precise measurement (number) used in the calculation.
- **Example:** A book 50.85 mm thick, a box 168.3 mm thick and a piece of paper 0.037 mm thick are stacked on top of each other. What is the height of the stack?

$$\begin{array}{r l}
50.85\text{ mm} & \\
\boxed{168.3\text{ mm}} & \leftarrow \text{Least precise: precise to tenths} \\
+\ 0.037\text{ mm} & \\
\hline
219.187\text{ mm} & \xrightarrow{\text{Round to tenths}} \boxed{219.2\text{ mm}}
\end{array}$$

Understanding Check: Each of the numbers below is measured. Solve the calculations and give the correct number of significant figures.

a) 0.12 x 1.77
b) 690.4 ÷ 12
c) 5.444 – 0.44
d) 16.5 + 0.114 + 3.55

Unit Conversions

Typical Unit Conversion Problems:

- A package weighs 3.50 kg (kilograms), what is the weight in lbs. (pounds)
- A student is 60.0 inches tall, what is the student's height in cm?
- The temperature in Cabo San Lucas, Mexico is 30^{o}C, what is the temperature in ^{o}F?

To convert from one unit to another, we must know the ___________________ between the two units of measure.

- Examples:
  - A package weighs 3.50 kg (kilograms), what is the weight in lbs. (pounds)
    – 1kg = 2.20 lb
  - A student is 60.0 inches tall, what is the student's height in cm?
    – 1 inch = 2.54 cm

Unit Relationships to Know:

- 1 milliliter (mL) = 1 cubic centimeter (cm^3)
- 1 inch (in) = 2.54 centimeters (cm)
- 1 kilogram (kg) = 2.20 pounds (lb)
- 4.184 Joule (J) = 1 calorie (cal)

The *relationships* between units are called ______________________ _______________.

Unit Conversion Calculations: The Factor Label Method

A package weighs 3.50 kg (kilograms), what is the weight in lbs. (pounds)?
Equivalence statement: 1kg = 2.20 lb

$$\left(\frac{1\text{ kg}}{2.20\text{ lb}}\right) \quad \text{Conversion Factors} \quad \left(\frac{2.20\text{ lb}}{1\text{ kg}}\right)$$

Equivalence statements can be written as ________________ ______________.

Take notes here:

=

Some conversion factors have an _________________ number of significant figures.

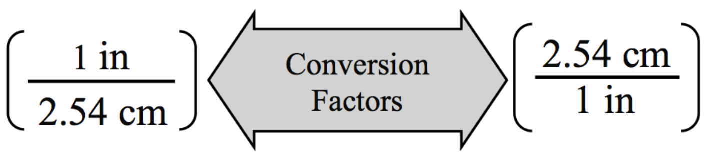
__________ (**defined or agreed upon**) conversion factors have an ***infinite*** number of significant figures.

•Examples of exact/defined conversion factors

- 1 lb = 0.45359237 kg
- 1 inch = 2.54 cm
- 1 cg = 10^{-2}g
- 1 ft = 12 inches
- 1 ml = 1cm^3

A student is 60.0 inches tall, what is the student's height in cm?

Equivalence statement: 1 inch = 2.54 cm



$$\underbrace{60.0}_{\text{three significant figures}}\,\cancel{\text{in}}\ \underbrace{\left(\frac{2.54\text{ cm}}{1\ \cancel{\text{in}}}\right)}_{\text{an }\textbf{infinite}\text{ number of significant figures}} = 152\text{ cm}$$

Take notes here:

=

Understanding Check:

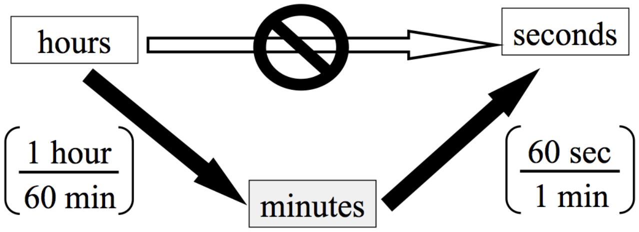
1) How many ft. (feet) in 379.3 in. (inches)?
   - 1 ft = 12 inches

2) How many eggs in 7.5 dozen?
   - 12 eggs = 1 dozen

3) How many calories in 514 joules?
   - 1 calorie = 4.184 joules

Sometimes it takes more than one step!

•Example: How many seconds in 33.0 hours?



Take notes here:

Take notes here:

Now you try a two-step conversion: How many inches in 5.5 meters given that:

- 1 inch = 2.54 cm
- 100 cm = 1 m

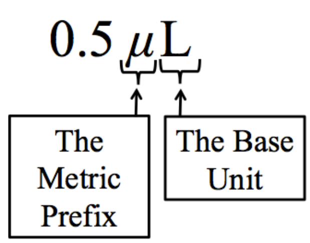
Metric Prefixes

Earlier, we used scientific notation to simplify working with very large or very small numbers.

Another way to simplify working with large or small numbers is to use metric _______________.

Example: The volume of blood required for diabetics to measure blood glucose levels in modern glucometers is about 0.0000005 L.

It is much more practical to use and say:



The metric **prefix** tells the ***fraction*** or ***multiple*** of the **base unit(s)**.

- For example, $1 \times 10^6\ \mu L = 1\ L$

The base unit can be ________________metric unit:

- liter (L), gram (g), meter (m), joule (J), second (s), calorie (cal)…etc.

Unit Conversions Within The Metric System

Example: The volume of blood required to measure blood glucose levels in modern glucometers is about 0.0000005 L.

- Question: How can we convert that to μL?
- Answer: We need the relationship between L and μL to get the conversion factor.

We will use the "Equality Table":

1 base unit =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

All these quantities in the table are equal; any pair can be used as a conversion factor!!!

Example: What is the relationship between L (microliters) and liters (L)?

1 base unit (Liters in this problem) =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

Equivalence statement: 1 L = 1 x 10^6 μL

This table works for **any** units!

The __________ ____________could be gram (g), meter (m), liter (L), joule (J), second (s), mole (mol), calorie (cal)… etc.

Understanding Check: Find the relationships between the following:

________ L = ________ mL

________ kg = ________ mg

________ nm = ________ m

________ cm = ________ mm

Example: How many μL (microliters) in 0.0000005 L?

$$\left(\frac{1 \times 10^6\ \mu L}{1\ L}\right) \longleftrightarrow \text{Conversion Factors} \longleftrightarrow \left(\frac{1\ L}{1 \times 10^6\ \mu L}\right)$$

Equivalence statement: 1 L = 1 x 10^6 μL

1 base unit (Liters in this problem) =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

Take notes here:

=

You try one: How many mL (milliliters) in 0.0345 (kL) kiloliters?

Equivalence statement: ______ mL = _______ kL

=

1 base unit (Liters in this problem) =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

You try another one: A vial contains 9758 mg of blood serum. Convert this into grams (g).

Equivalence statement: ____________ g = ____________ mg

Temperature Unit Conversions

$$°F = (1.8 \times °C) + 32$$

$$°C = \frac{(°F - 32)}{1.8}$$

$$K = °C + 273.15$$

- Note: The 273.15, 32, and 1.8 in the temperature conversion equations are *exact.*

When doing a calculation that involves **only** multiplication and/or division, you can do the entire calculation then round the answer to the correct number of significant figures at the end. The same is true for a calculation that involves **only** addition and/or subtraction. But what about a calculation that involves mixed operations: **both** multiplication or division ***and*** addition or subtraction?

When doing calculations that involve **both** multiplication or division ***and*** addition or subtraction, first do a calculation for the operation *shown in parenthesis* and round that value to the correct number of significant figures, **then** use the rounded number to carry out the next operation.

Example: On a warm summer day, the temperature reaches 85 °F. What is this temperature in °C?

The relationship between ^{o}F and ^{o}C is:

$$°C = \frac{(°F - 32)}{1.8}$$

First, we do the subtraction (operation in parenthesis) and round the calculated value to the correct number of significant figures based on the rule for addition/subtraction.

Next, we divide that rounded number by 1.8 (exactly 1.8 = 1.80000....) then round the calculated value to the correct number of significant figures using the rule for multiplication/division.

Take notes here:

Significant Figures, Scientific Notation, and Rounding

1) Determine the number of significant figures in the following values:

Value	**# of sig. figures**	**Value**	**# of sig. figures**
140.74		4	
0.0041		3.70×10^{14}	
31.00		1.05×10^{12}	
1300		7.0400×10^{3}	
847.040		2495	

2) Round the following values to 3 significant figures.

3.76411 →	0.0411984 →
3.76811 →	150.6142 →
3.76511 →	0.013877 →
11.048176 →	4.88223×10^{9} →
8.75510 →	2.0097×10^{-12} →

3) Perform the following calculations and round the final answer to the correct number of significant figures.

Calculation	Rounded Answer	Calculation	Rounded Answer
18.7644 - 3.472 + 0.4101	=	0.87 + 4.061 + 10.4	=
17.441÷ 3	=	16 x 841.1 ÷ 16.300	=
14.044 + 8.11 + 3.4	=	21.01 x 2.0	=
3.41 – 0.086652	=	18.4 +12.99 +13.772 + 9.704	=

4. Convert the following into scientific notation or standard notation

Standard notation	**Scientific notation**
47,000	
0.0008	
675,000,000	
157,000,000,000,000,000,000,000	
0.0000003407	
	7.66×10^{-2}
	7.8×10^{5}
	4.75×10^{-4}
	6×10^{-3}
	9×10^{8}

Conversions Within the Metric System:
Perform the following metric conversions. Show your conversion factors. Use correct number of significant figures. If you need more room, do calculations on separate page(s.

0.50 m = ___________mm	2.00 km = ___________m	0.4000 L = __________mL
1.00 g = ___________kg	01.00 cm = __________m	8.00 mm = __________cm
22.4 L = ___________mL	5.00 g = ___________kg	4.245 L = __________ mL
345 g = ___________kg	10.0 nm = ___________m	3.22 Gg = ___________kg
3.001 cg = _________mg	1.2 m = ___________µm	455 nm = ___________m

English-Metric Conversions **(show your work)**

10.0 cm = ___________in	15.0 lb = ___________kg (1kg=2.205 lb)
1.00 yd = ___________ cm (1 yard = exactly 36 in)	16.9 fl. oz = ___________L (0.0338 fl oz.= 1 mL)
1.00 qt = ___________ L(1qt = 946 mL)	6.00 in = ___________cm
0.800 kg = ___________oz (16 oz = exactly 1 lb and 1kg=2.205 lb)	1.83 kg = ___________lb (1kg=2.205 lb)
25.00 mL = ___________qt (1qt = .946L)	1.40 L = ___________ = cm^3 note: 1 mL = exactly $1cm^3$

<u>Temperature Conversions</u>

Recall the Temperature Conversions from Chapter 1 lecture notes:

- $°F = (1.8 \times °C) + 32$
- $°C = (°F - 32) / 1.8$
- $K = °C + 273.15$

NOTE: In temperature conversion equations, the 273.15, 32 and 1.8 are ***exact.***

IMPORTANT: When doing a calculation that involves **only** multiplication and/or division, you can do the entire calculation then round the answer to the correct number of significant figures at the end. The same is true for a calculation that involves **only** addition and/or subtraction.

But what about a calculation that involves mixed operations: **both** multiplication or division ***and*** addition or subtraction?

When doing calculations that involve **<u>both</u>** multiplication or division ***and*** addition or subtraction, first do a calculation for the operation *shown in parenthesis* and round that value to the correct number of significant figures, **then** use the rounded number to carry out the next operation.

Perform the following temperature conversions (show your calculation)

75^oC = _________ K
-15^oC = _________ K
0.00 K = ______ ^oC = ______ ^oF
25^oC (room temperature) = _________ K
98.6 ^oF (body temperature) = _________ ^oC
25^oC = _________ ^oF
-40.0 ^oC = _________ ^oF
412 K = _________ ^oF

Significant Figures, Scientific Notation, and Rounding

1) Determine the number of significant figures in the following values:

Value	**# of sig. figures**	**Value**	**# of sig. figures**
140.74	5	4	1
0.0041	2	3.70×10^{14}	3
31.00	4	1.05×10^{12}	3
1300	2	7.0400×10^{3}	5
847.040	6	2495	4

2) Round the following values to 3 significant figures.

3.76411 →3.76	0.0411984 →0.0412
3.76811 →3.77	150.6142 → 151
3.76511 →3.77	0.013877 → 0.0139
11.048176 →11.0	$4.88223 \times 10^{9} \rightarrow 4.88 \times 10^{9}$
8.75510 →8.76	$2.0097 \times 10^{-12} \rightarrow 2.01 \times 10^{-12}$

3) Perform the following calculations and round the final answer to the correct number of significant figures.

Calculation	Rounded Answer	Calculation	Rounded Answer
18.7644 - 3.472 + 0.4101	= 15.703	0.87 + 4.061 + 10.4	= 15.3
17.441÷ 3	= 6	16 x 841.1 ÷ 16.300	= 830
14.044 + 8.11 + 3.4	= 25.6	21.01 x 2.0	= 42
3.41 – 0.086652	= 3.32	18.4 +12.99 +13.772 + 9.704	= 54.9

4. Convert the following into scientific notation or standard notation

Standard notation	**Scientific notation**
47,000	4.7×10^{4}
0.0008	8×10^{-4}
675,000,000	6.75×10^{8}
157,000,000,000,000,000,000,000	1.57×10^{23}
0.0000003407	3.407×10^{-7}
0.0766	7.66×10^{-2}
780,000	7.8×10^{5}
0.000475	4.75×10^{-4}
0.006	6×10^{-3}
900,000,000	9×10^{8}

Metric System:

Perform the following metric conversions. Show your conversion factors. Use correct number of significant figures. If you need more room, do calculations on separate page(s).

0.50 m = _5.0 x 10^2___mm	2.00 km = __2.00 x 10^3__m	0.4000 L = 4.000 x 10^2 mL or 400.0 mL
1.00 g = _1.00 x 10^{-3}_kg or .00100 kg	01.00 cm = __0.0100__m or 1.00 x 10^{-2} m	8.00 mm = __0.800___cm or 8.00 x 10^{-1} cm
22.4 L = 2.24 x 10^4 mL or 22400 mL	5.00 g = 5.00 x 10^{-3}_kg or .00500 kg	4.245 L = __4245__ mL or 4.245 x 10^3 mL
345 g = _0.345_kg or 3.45 x 10-^{1} kg	10.0 nm = 1.00 x 10^{-8}_m or .0000000100 m	3.22 Gg = 3.22x 10^6 kg or 3220000 kg
3.001 cg = _30.01_mg or 3.001 x 10^1 mg	1.2 m = _1.2 x 10^6_µm or 1200000 µm	455 nm = _4.55 x 10^{-7}_m or .000000455 m

***English-Metric Conversions* (show your work)**

10.0 cm = 3.94_in	15.0 lb = ___6.80__kg
1.00 yd = __91.4____ cm	16.9 fl. oz = __0.500___L (0.0338 fl oz.= 1 mL)
1.00 qt = ___0.946___ L	6.00 in = ___15.2___cm
0.800 kg = ___28.2___oz (16 oz = 1 lb)	1.83 kg = ___4.04___lb
25.00 mL = ___0.0264___qt (1qt = .946L)	1.40 L = __1.40 x 10^3___ = cm^3 note: 1 mL = $1cm^3$

Temperature Conversions

Perform the following temperature conversions (show your calculation)

75^{o}C = __348___ K
-15^{o}C = ___258___ K
0.00 K = __-273.15_ ^{o}C = __-459.67__ ^{o}F
25^{o}C (room temperature) = ____298___ K
98.6 ^{o}F (body temperature) = ___37.0___ ^{o}C
25^{o}C = ___77____ ^{o}F
-40.0 ^{o}C = __-40.0___ ^{o}F
412 K = __282____ ^{o}F

Chapter 1 Worksheet 1 and KEY

Significant Figures Worksheet

1. Indicate how many significant figures there are in each of the following measured values.

246.32	________	1.008	________	700000	________
107.854	________	0.00340	________	350.670	________
100.3	________	14.600	________	1.0000	________
0.678	________	0.0001	________	320001	________

2. Calculate the answers to the appropriate number of significant figures.

32.567	246.24	658.0
135.0	238.278	23.5478
+ 1.4567	+ 98.3	+ 1345.29

3. Calculate the answers to the appropriate number of significant figures.

a) 23.7 x 3.8 = ________

b) 45.76 x 0.25 = ________

c) 81.04 x 0.010 = ________

d) 6.47 x 64.5 = ________

e) 43.678 x 64.1 = ________

f) 1.678 / 0.42 = ________

g) 28.367 / 3.74 = ________

h) 4278 / 1.006 = ________

Significant Figures Worksheet Key

1. Indicate how many significant figures there are in each of the following measured values.

246.32	5 sig figs	1.008	4 sig figs	700000	1 sig fig
107.854	6 sig figs	0.00340	3 sig figs	350.670	6 sig figs
100.3	4 sig figs	14.600	5 sig figs	1.0000	5 sig figs
0.678	3 sig figs	0.0001	1 sig fig	320001	6 sig figs

2. Calculate the answers to the appropriate number of significant figures.

32.567	246.24	658.0
135.0	238.278	23.5478
+ 1.4567	+ 98.3	+ 1345.29
169.0	**582.8**	**2026.8**

3. Calculate the answers to the appropriate number of significant figures.

a) 23.7 x 3.8 = **90.**

b) 45.76 x 0.25 = **11**

c) 81.04 g x0.010 = **0.81**

d) 6.47 x 64.5 = **417**

e) 43.678 x 64.1 = **2.80×10^3**

f) 1.678 / 0.42 = **4.0**

g) 28.367 / 3.74 = **7.58**

h) 4278 / 1.006 = **4252**

Unit Conversions Worksheet

Complete each of the following conversions to the proper number of significant figures; clearly show your set-up with units in the set-up and the answer.

1) 0.30 m = ________________ mm

2) 5.00 mm = ________________ cm

3) 10.0 cm = ________________ in

4) 2.00 km = ________________ m

5) 33.4 L = ________________ mL

6) 15.0 lb = ________________ kg
(use 2.20 lb = 1 kg)

7) 0.400 L = ________________ mL

8) 5.00 g = ________________ kg

9) 1.00 yd = ________________ cm
(3 ft = 1 yard, exactly)

10) 6.35 g = ________________ kg

11) 8.245 L = ________________ mL

12) 16.9 fl. oz= ________________ L
(0.0338 fl oz.= 1 mL)

13) 1.00 cm = ________________ m

14) 345 g = ________________ kg

See next page for Key

Unit Conversion Worksheet Key

1) 0.30 m = 3.0×10^2 mm

2) 5.00 mm = 0.500 cm <u>or</u> 5.00×10^{-1} cm

3) 10.0 cm = 3.94 in

4) 2.00 km = 2.00×10^3 m

5) 33.4 L = 3.34×10^4 mL <u>or</u> 33400 mL

6) 15.0 lb = 6.82 kg

7) 0.400 L = 400. mL <u>or</u> 4.00×10^2 mL

8) 5.00 g = 5.00×10^{-3} kg <u>or</u> 0.00500 kg

9) 1.00 yd = 91.4 cm

10) 6.35 g = 0.00635 kg <u>or</u> 6.35×10^{-3} kg

11) 8.245 L = 8245 mL <u>or</u> 8.245×10^3 mL

12) 16.9 fl. oz = 0.500 L <u>or</u> 5.00×10^{-1} L

13) 1.00 cm = 0.0100 m <u>or</u> 1.00×10^{-2} m

14) 345 g = 0.345 kg <u>or</u> 3.45×10^{-1} kg

Chapter 2 Lecture Notes: Atoms

Educational Goals

1. Describe the subatomic structure of an atom.
2. Define the terms **element** and **atomic symbol.**
3. Understand how elements are arranged in the periodic table based on the number of protons they contain.
4. Understand how **atomic number** and **mass number** are used to indicate details of an atom's nucleus.
5. Know how **isotopes** of an element differ from one another.
6. Define the term **mole** and describe the relationship between **moles** and **molar mass**.
7. Given the **molar mass** of an element, convert between number of atoms, number of moles, and mass (grams).

An Introduction to Atoms

Matter (stuff) is made of _______________.

Model of the Atom

Check your current model: Draw a carbon atom.

Atoms are made of _________________ ***particles***.

There are *three* types of subatomic particles that will make up our atomic model:

1. ______________
2. ______________
3. ______________

Protons and neutrons are compacted together in what we call the ________________ of an atom. Electrons are distributed in space around the nucleus.

- They are moving very fast in a volume surrounding the nucleus.

Atoms are mostly empty space.

Electrical Charge

There are a few fundamental properties of nature.

- Examples: Gravity, magnetism, and mass.

Another fundamental property in nature is ________________ ______________.

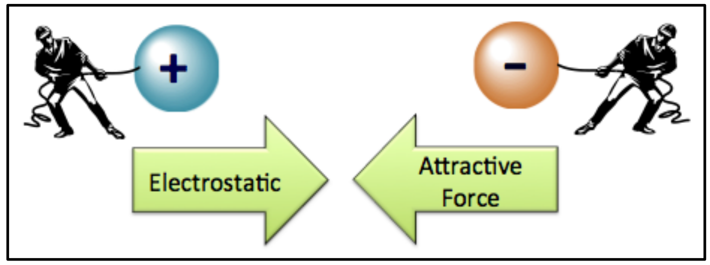
Particles ***may or may not*** have electrical charge.

There are two types of electrical charge; we arbitrarily call one type ______________ and the other type ___________.

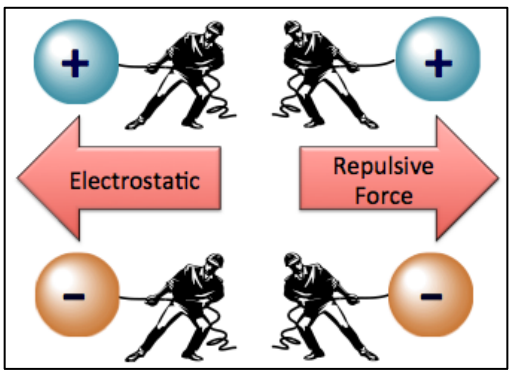
Every thing we discuss in this course ultimately occurs because of the interaction of these two types of charges.

Particles with ***opposite charges attract*** each other.

The natural attraction is called __________________ __________________ **force**.



Oppositely charged particles will accelerate **toward** one another if not held apart.



Particles with ***like charges repel*** each other.

The natural repulsion is called __________________ __________________ **force**.

Like charged particles will accelerate **away** from one another if not held together.

Subatomic Particles

1) Protons

Protons are ________________ ***charged*** particles located in the _______________ of an atom.

The number of protons a particular atom contains determines that atom's identity.

- For example, any atom that contains just **one proton** is called *hydrogen*. An atom with **two protons** is called *helium*. An atom with **six protons** is called *carbon*.

Historically, matter with different numbers of protons, such as hydrogen, helium, and carbon were called the _________________.

There are 92 elements that occur in nature. About 25 others have been man-made by slamming two atoms together causing their nuclei to combine, however these new atoms do not last long (fractions of a second up to one year), they break apart into smaller atoms.

A modern periodic table of the elements is shown on the next page.

- You can download a copy of this periodic table at: http://www.zovallearning.com/GOBlinks/ch2/periodictablezovalbasic.pdf

Periodic Table of the Elements

I Alkali Metals	II Alkaline Earth Metals											III	IV	V	VI	VII Halogens	VIII Noble Gases
1 **H** Hydrogen 1.0079																	2 **He** Helium 4.003
3 **Li** Lithium 6.941	4 **Be** Beryllium 9.012183											5 **B** Boron 10.811	6 **C** Carbon 12.0107	7 **N** Nitrogen 14.0067	8 **O** Oxygen 15.9994	9 **F** Fluorine 18.998403	10 **Ne** Neon 20.1797
11 **Na** Sodium 22.989770	12 **Mg** Magnesium 24.3050											13 **Al** Aluminum 26.9815	14 **Si** Silicon 28.0855	15 **P** Phosphorus 30.973762	16 **S** Sulfur 32.066	17 **Cl** Chlorine 35.4527	18 **Ar** Argon 39.948
19 **K** Potassium 39.0983	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.955908	22 **Ti** Titanium 47.867	23 **V** Vanadium 50.9415	24 **Cr** Chromium 51.9961	25 **Mn** Manganese 54.938044	26 **Fe** Iron 55.845	27 **Co** Cobalt 58.933194	28 **Ni** Nickel 58.6934	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92160	34 **Se** Selenium 78.971	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.80
37 **Rb** Rubidium 85.4678	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.90584	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.90637	42 **Mo** Molybdenum 95.95	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.90550	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.8682	48 **Cd** Cadmium 112.414	49 **In** Indium 114.818	50 **Sn** Tin 118.710	51 **Sb** Antimony 121.760	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90447	54 **Xe** Xenon 131.29
55 **Cs** Cesium 132.90545	56 **Ba** Barium 137.327	57 **La** Lanthanum 138.90545	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.9479	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.207	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.217	78 **Pt** Platinum 195.078	79 **Au** Gold 196.96657	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.3833	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.98038	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **Ac** Actinium (227)	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (263)	107 **Bh** Bohrium (262)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (266)									

58 **Ce** Cerium 140.116	59 **Pr** Praseodymium 140.90766	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.964	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.92534	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93033	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.93422	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.967
90 **Th** Thorium 232.0377	91 **Pa** Protactinium 231.03588	92 **U** Uranium 238.0289	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Note that each element is represented by its **atomic** ______________ (a one- or two-letter name abbreviation) and occupies a box in the table.

Above each element's symbol is the ______________ ________________.

The **atomic number** tells us the _____________ ***of*** ____________ in an atom of that particular element.

- Example: Look at carbon, symbol C, atomic number 6. Carbon has an atomic number of *six* because an atom with six protons is called carbon. If it had *seven* protons, it would not be carbon - it would be nitrogen and have an atomic number of 7.
- Atomic number can be abbreviated using "**Z**."
  - For example, with carbon, **Z** = 6, with hydrogen, **Z** = 1.
- Elements are ordered in the periodic table by *increasing* atomic number.

2) Electrons

Electrons are ***negatively charged*** subatomic particles.

They are light-weight particles that move extremely fast.

- For the remainder of chapter 2 we can visualize the electrons as bees flying around a beehive (the bee hive represents the nucleus). In chapter 3 you will learn more details about the regions around the nucleus that the electrons can occupy.
- Electrons are very light compared to protons and neutrons.
- Protons and neutrons are about 2000 times **heavier** than electrons and therefore compose most of an atom's mass.

3) Neutrons

Neutrons are located in the _______________ (with the protons).

Neutrons **do not** have electrical charge; we say they are ***electrically*** __________________.

The names, charges, and symbols for the three types of subatomic particles are shown below:

SUBATOMIC PARTICLE	SYMBOL	CHARGE
PROTON	*p*	positive (1+)
NEUTRON	*n*	none
ELECTRON	*e* or e^-	negative (1-)

How many neutrons are in an atom?

We ***cannot determine*** the number of neutrons in an atom based on the number of protons.

- This is because atoms of a particular element ***do not all have the same number of neutrons***.

Example: Some carbon atoms have *six neutrons*, some have *seven neutrons*, and some have *eight neutrons*.

- These three different forms of carbon are called _____________ of carbon.

Isotopes are defined as atoms with the *same* number of protons (same element), but a *different* number of neutrons.

You learned that an atom's *"atomic number* (Z)*"* is the *number of protons* it contains.

When considering the number of neutrons in an isotope of a particular atom, it is useful to learn a new term called "**mass number**."

The _________ ____________ of an atom is defined as *the number of protons plus the number of neutrons*.

mass number = *number of* ***protons*** + *number of* ***neutrons***

Mass number can be abbreviated using "A."

	SYMBOL	DEFINITION
ATOMIC NUMBER	Z	number of protons
MASS NUMBER	A	number of protons + number of neutrons

Example: How many neutrons are in a sodium (Na) atom that has a ***mass number*** of **23**?

Take notes here:

Understanding Check: How many neutrons are in a **carbon** (**C**) atom that has a ***mass number*** of **14**?

You will often see one of two "shorthand notation" methods used to differentiate the various isotopes:

Method 1: Write the *element symbol*, a dash, then the *mass number* (A)

Let's use our three isotopes of carbon for examples:

NUMBER OF NEUTRONS IN THE CARBON ATOM	SHORTHAND NOTATION
6	C-12
7	C-13
8	C-14

Method 2: Write the *element symbol*, we superscript the *mass number* (A) to the left of the symbol.

NUMBER OF NEUTRONS IN THE CARBON ATOM	SHORTHAND NOTATION
6	^{12}C
7	^{13}C
8	^{14}C

- Although redundant, sometimes the atomic number (Z) is also subscripted to the left of the symbol.
  - For example:

$$^{12}_{6}C$$

Understanding Check: Fill in the blanks for the following isotopes:

a. **^{14}N** number of protons ___ number of neutrons ___ atomic number ___ mass number ___
b. **^{15}N** number of protons ___ number of neutrons ___ atomic number ___ mass number ___
c. **^{42}Ca** number of protons ___ number of neutrons ___ atomic number ___ mass number ___
d. **^{1}H** number of protons number of neutrons atomic number mass number

Atoms are ***electrically neutral***; their **total charge** is equal to **zero**.

- They have the same number of electrons (-) as protons (+), so the positive and negative charges add up to zero (cancel).

The Mole

Atoms are so tiny and small in mass that it is more convenient to do calculations with a large number of atoms

- Just like bakers and chefs use eggs by the dozen, chemists use atoms and molecules by the mole.
  - A ____________________ is a counting unit used for atoms and molecules.
  - A _______________ ______________is any term that refers to a specific number of things.
    - a couple = 2 items (e.g. people)
    - a dozen = 12 items (e.g. eggs, donuts)
    - a mole = 6.022×10^{23} (e.g. atoms, molecules)

The Chemist's Mole

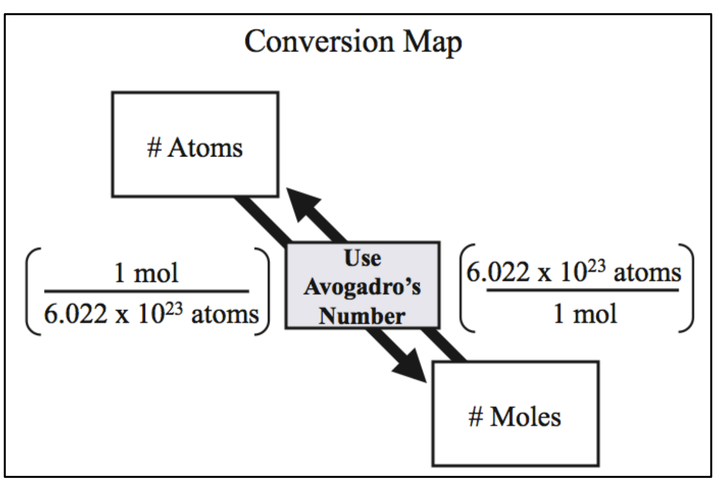
- **One mole** of anything represents 6.022×10^{23} of the things.
- This is referred to as **Avogadro's number**.
- 1 mole = 6.022×10^{23}

Understanding Check: How many atoms are in ***1 mole*** of helium (He)? ___________________________

Because the mole is the standard counting unit used to indicate the number of atoms present in a sample, it is useful to **convert** back and forth from ***moles*** to ***atoms***.

- Use our ***conversion factor*** method.
- The ***relationship*** between # of atoms and moles is:
  - 1 mole = 6.022×10^{23}

$$\left(\frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}\right) \quad \text{Conversion Factors} \quad \left(\frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}\right)$$



Example: How many carbon atoms are there in 0.100 mole of carbon?

Take notes here:

You try one: How many moles are 2.9×10^{12} F atoms?

The Mole and Mass

- The _______________ ______________ of an element is equivalent to the mass (in grams) of one mole of the element.
- Molar mass is given in the ***periodic table*** _______________ the symbol of the element.
  - Molar mass units: ___________________
  - Example: Carbon – molar mass is _________________
  - Another example:
    - 1 mole of argon (Ar) = 39.95 g
    - Molar mass of argon is 39.95 g/mole

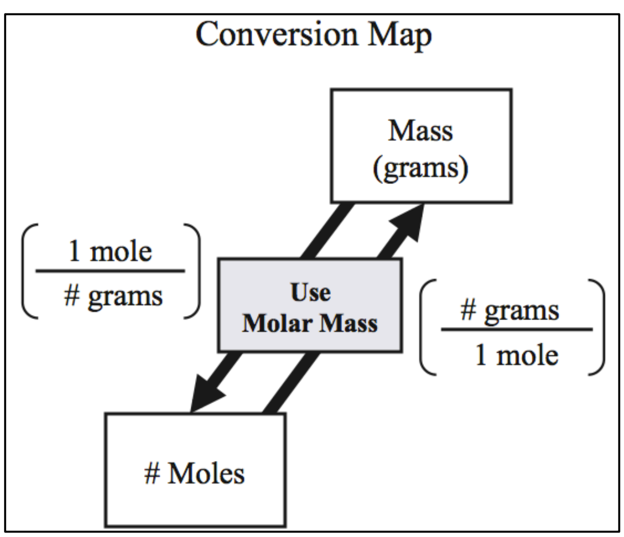
Understanding Check:

1 mole of C = _____ grams of carbon (C) = ____________ atoms of C

1 mole of Al = ______grams of aluminum (Al) = _____________ atoms of Al

Because the molar mass gives us the ______________________between the number of moles and the mass of an element, it can be used to _____________ back and forth between moles and mass (in grams).

- Use our conversion factor method



Example: **Carbon**

- The relationship between # of moles of carbon and grams of carbon is:
  - 1 mole Carbon = 12.01 g
- This can be written as conversion factors:

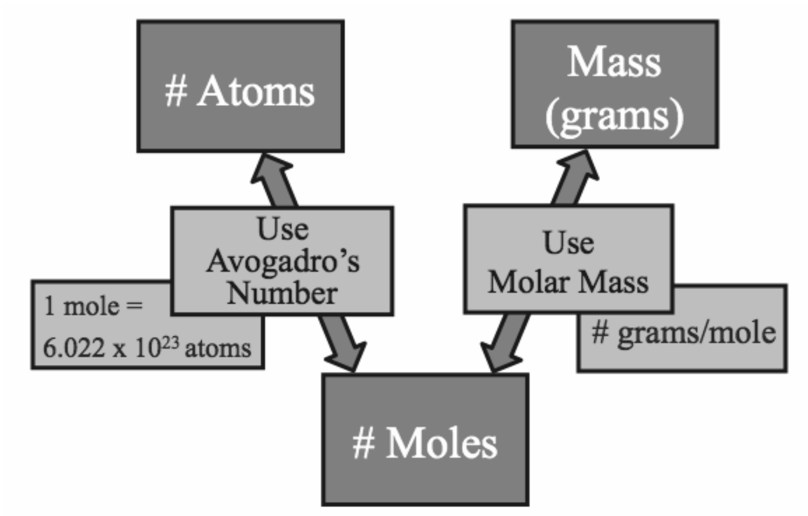
$$\left(\frac{1 \text{ mole C}}{12.01 \text{ grams C}}\right) \quad \text{Conversion Factors} \quad \left(\frac{12.01 \text{ grams C}}{1 \text{ mole C}}\right)$$

Example Problem: What is the mass of 0.770 moles of carbon?

Take notes here:

You try one: How many moles are there in 50.0 g of lead?

Converting Between the Number of Atoms and Grams



Example: (atoms to grams) What is the mass of 2.50×10^{21} Lead (Pb) atoms?

Take notes here:

You try one: (grams to atoms) Compute the number atoms in 10.0 g of Aluminum (Al)?

The Periodic Table

As we continue to build our model of atoms and matter in later chapters, we will gain more understanding of why the elements are arranged as they are in the periodic table and how the periodic table can be very useful in predicting the chemical and physical properties of matter.

Classification of Elements Based on Electrical and Heat Conduction

CATEGORY	PROPERTIES
Metals	•Good conductors of heat and electricity •Ductile (can be pulled into wires and pounded flat) •Have a luster
Nonmetals	•Poor conductors of heat and electricity •Brittle (break or shatter if bent or hammered)
Metalloids (sometimes called Semimetals)	Intermediate conductors of heat and electricity

1 **H**			Metals (Green)			Nonmetals (Blue)			Metalloids (Red)								2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**									
				58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
				90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

Elements in the periodic table are arranged in columns called ____________ (sometimes, but much less often, called **Families)**.

- Sometimes these groups are shown with **group numbers** in Roman numerals above the column.

s-Block p-Block

d-Block f-Block

Transition Metals

	I	II											III	IV	V	VI	VII	VIII
1	1 **H**																	2 **He**
2	3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
3	11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
4	19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
5	37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
6	55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
7	87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**									

(Inner) Transition Metals

6	Lanthanides	58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
7	Actinides	90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

The elements in **Group I** (also called Group 1A) are called the ____________ ***metals***.

- Although it is not a metal, note that hydrogen is in this group *for reasons that I will discuss in chapter 3*.

The elements in **Group II** (also called group 2A) are called the ____________ ***earth metals***.

The elements in **Group VII** (also called group 7A) are called the ____________.

The elements in **Group VIII** (also called group 8A) are called the ____________**.**

The elements in **Group I** and **Group II** are in what is called the ____ **-Block**.

The elements in Groups **III - VIII** are in the ____ **-Block**.

The _________________ ______________, *located between the s- and p-Blocks*, are in the ___ **-Block**.

The **Inner Transition Metals**, located in the bottom two rows of the periodic table are in the ___ **-Block**.

- They are called ***lanthanides*** (top row of the *f-Block*) and ***actinides*** (bottom row of the *f-Block*).

The ***rows*** in the periodic table are called _______________.

- The periods are often numbered to the left of each row.

Chapter 2: Atomic Molar Mass Worksheet and Key

1. Complete the following table:

Symbol	Symbol-Mass Number	Atomic #	Mass #	# protons	# neutrons	# electrons
$^{7}_{3}Li$						
	Mo-96					
				49	53	
			72			35
$^{238}_{92}U$						

Answer the following questions. ***Be sure to write units with every number and to use the correct number of significant figures***. Use **two digits to the right of the decimal place for molar masses when possible**, then your final answer will match the key exactly.

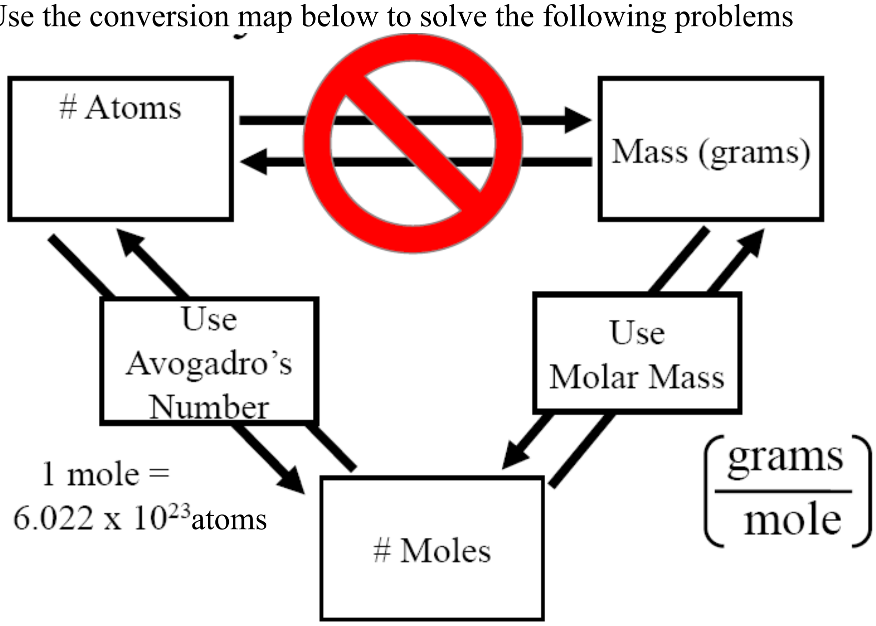
1. What is the molar mass of the following elements?

a) B _______________ b) Zn _______________ c) He __________________

2. What is the mass (grams) of one mole of Xenon?____

3. How many atoms are in one mole of Xenon? ____

Use the conversion map below to solve the following problems



4. How many moles of oxygen (O) contain 1.65×10^{27} oxygen atoms?

1.65×10^{27} O atoms	_____ mole O	=	________ mole O
	O atoms		

5. How many moles of K contain 3.50×10^{19} K atoms?

3.50×10^{19} K atoms		=	_________ mole K

6. How many moles of fluorine (F) contain 8.27×10^{24} F atoms?

		=

7. How many carbon atoms are contained in 5.45 moles of carbon?

5.45 moles C	__________ C atoms	=	________________ C atoms
	___ mole C		

8. How many Neon (Ne) atoms are contained in 0.75 moles of Ne?

		=	_____

9. How many moles of oxygen (O) is contained 2.65 grams of oxygen?

2.65 g O	1 mole O	=	_____ mole O
	_______ g O		

10. How many moles of potassium (K) is contained 8.44 grams of potassium?

8.44 g K		=	_____ mole K

11. How many g of Xe is contained in 0.054 moles of Xe?

0.054 moles Xe	________ g Xe	=	_____
	___ mole Xe		

12. How many g of C is contained in 39.5 moles of C?

		= ____

13. What is the mass (grams) of 5.00×10^{24} oxygen atoms?

5.00×10^{24} atoms O	________ mole O	______ g O	= _______ g O
	________________atoms O	________ mole O	

14. What is the mass (grams) of 1.00×10^{6} sodium (Na) atoms?

1.00×10^{6} Na atoms			= _________ g Na

15. How many iron atoms are in 10.0 grams of iron (Fe) ?

10.0 g Fe	________ mole Fe	____________ Fe atoms	= _____________Fe atoms
	_______ g Fe	________ mole Fe	

16. How many copper (Cu) atoms are in a 257 gram copper pipe?

			= _____________Cu atoms

See next page for KEY

KEY

1. Complete the following table:

Symbol	Symbol- Mass Number	Atomic #	Mass #	# protons	# neutrons	# electrons
$^{7}_{3}Li$	Li-7	3	7	3	4	3
$^{96}_{42}Mo$	Mo-96	42	96	42	54	42
$^{102}_{49}In$	In-102	49	102	49	53	49
$^{72}_{35}Br$	Br-72	35	72	35	37	35
$^{238}_{92}U$	U-238	92	238	92	146	92

Answer the following questions. ***Be sure to write units with every number and to use the correct number of significant figures***. Use **two digits to the right of the decimal place for molar masses when possible**, then your final answer will match the key exactly.

1. What is the molar mass of the following elements?

a) B 10.81g/mole b) Zn 65.39 g/mole c) He 4.00 g/mole

2. What is the mass (grams) of one mole of Xenon? 131.29 g

3. How many atoms are in one mole of Xenon? 6.022×10^{23}

4. How many moles of oxygen (O) contain 1.65×10^{27} oxygen atoms?

$$1.65 \times 10^{27} \text{ O atoms} \times \frac{1 \text{ mole O}}{6.022 \times 10^{23} \text{ O atoms}} = 2740 \text{ mole O } \textbf{or } 2.74 \times 10^{3} \text{ mole O}$$

Note: 3 significant figures!!!

5. How many moles of K contain 3.50×10^{19} K atoms?

$$3.50 \times 10^{19} \text{ K atoms} \times \frac{1 \text{ mole K}}{6.022 \times 10^{23} \text{ K atoms}} = 5.81 \times 10^{-5} \text{ mole K}$$

6. How many moles of fluorine (F) contain 8.27×10^{24} F atoms?

$$8.27 \times 10^{24} \text{ F atoms} \times \frac{1 \text{ mole F}}{6.022 \times 10^{23} \text{ F atoms}} = 13.7 \text{ mole F}$$

7. How many carbon atoms are contained in 5.45 moles of carbon?

$$\frac{5.45 \text{ moles C}}{} \times \frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mole C}} = 3.28 \times 10^{24} \text{ C atoms}$$

8. How many Neon (Ne) atoms are contained in 0.75 moles of Ne?

$$\frac{0.75 \text{ moles Ne}}{} \times \frac{6.022 \times 10^{23} \text{ Ne atoms}}{1 \text{ mole Ne}} = 4.5 \times 10^{23} \text{ Ne atoms}$$

9. How many moles of oxygen (O) is contained 2.65 grams of oxygen?

$$\frac{2.65 \text{ g O}}{} \times \frac{1 \text{ mole O}}{16.00 \text{ g O}} = 0.166 \text{ mole O}$$

Molar Mass of Oxygen
1 mole O = 16.00 grams O

10. How many moles of potassium (K) is contained 8.44 grams of potassium?

$$\frac{8.44 \text{ g K}}{} \times \frac{1 \text{ mole K}}{39.10 \text{ g K}} = 0.216 \text{ mole K}$$

11. How many g of Xe is contained in 0.054 moles of Xe?

$$\frac{0.054 \text{ moles Xe}}{} \times \frac{131.29 \text{ g Xe}}{1 \text{ mole Xe}} = 7.1 \text{ g Xe}$$ (NOTE: 2 significant figures)

12. How many g of C is contained in 39.5 moles of C?

$$\frac{39.5 \text{ mole C}}{} \times \frac{12.01 \text{ g C}}{1 \text{ mole C}} = 474 \text{ g C}$$

13. What is the mass (grams) of 5.00 x 10^{24} oxygen atoms?

$$\frac{5.00 \times 10^{24} \text{ atoms O}}{} \times \frac{1 \text{ mole O}}{6.022 \times 10^{23} \text{ O atoms}} \times \frac{16.00 \text{ g O}}{1 \text{ mole O}} = 133 \text{ g O}$$

14. What is the mass (grams) of 1.00×10^6 sodium (Na) atoms?

1.00×10^6 Na atoms	1 mole Na	22.99 g Na
	6.022×10^{23} Na atoms	1 mole Na

$= 3.82 \times 10^{-17}$ g Na

15. How many iron atoms are in 10.0 grams of iron (Fe) ?

10.0 g Fe	1 mole Fe	6.022×10^{23} Fe atoms
	55.85 g Fe	1 mole Fe

$= 1.08 \times 10^{23}$ Fe atoms

16. How many copper (Cu) atoms are in a 257 gram copper pipe?

257 g Cu	1 mole Cu	6.022×10^{23} Cu atoms
	63.55 g Cu	1 mole Cu

$= 2.44 \times 10^{24}$ Cu atoms

CHAPTER 2 REVIEW WORKSHEET AND KEY

The Mole

1) How many zinc (Zn) atoms are contained in 5.16 moles of Zn?

2) How many moles of He are 221,000 He atoms?

3) How many atoms are contained in 0.98 moles of iron (Fe)?

4) How many moles of cesium are in 66.45 g Cs?

5) What is the mass (grams) of 2500. carbon atoms?

6) What is the mass (grams) of 6.52 x 10^{18} atoms of gold (Au)?

Isotopes

(IMPORTANT NOTE: X = Symbol, A=mass number, Z = atomic number)

Name	Atomic number	Mass number	# of protons	# of neutrons	X-A form	$^{A}_{Z}X$ notation
Cobalt		60				
					I-131	
						$^{3}_{1}H$
	26	59				
						$^{99}_{42}Mo$
			11	24		
Strontium				52		
					U-235	
		134	55			
		19	9			
	79			118		
Copper				36		
			56	81		
					K-40	

KEY

The Mole

1) How many zinc (Zn) atoms in 5.16 moles of Zn?

3.11×10^{24} Zn atoms

2) How many moles of He are 221,000 He atoms?

3.67×10^{-19} moles He

3) How many atoms are contained in 0.98 moles of iron (Fe)?

5.9×10^{23} Fe atoms

4) How many moles of cesium are in 66.45 g Cs?

0.5000 moles Cs

5) What is the mass (grams) of 2500. carbon atoms?

4.986×10^{-20}g C

6) What is the mass (grams) of 6.52×10^{18} atoms of gold (Au)?

0.00213 g Au

Isotopes

Name	**Atomic number**	**Mass number**	**# of protons**	**# of neutrons**	**X-A form**	$^{A}_{Z}X$ **notation**
Cobalt	27	60	27	33	Co-60	$^{60}_{27}Co$
Iodine	53	131	53	78	I-131	$^{131}_{53}I$
Hydrogen	1	3	1	2	H-3	$^{3}_{1}H$
Iron	26	59	26	33	Fe-59	$^{59}_{26}Fe$
Molybdenum	42	99	42	57	Mo-99	$^{99}_{42}Mo$
Sodium	11	35	11	24	Na-35	$^{35}_{11}Na$
Strontium	38	90	38	52	Sr-90	$^{90}_{38}Sr$
Uranium	92	235	92	143	U-235	$^{235}_{92}U$
Cesium	55	134	55	79	Cs-134	$^{134}_{55}Cs$
Fluorine	9	19	9	10	F-19	$^{19}_{9}F$
Gold	79	197	79	118	Au-197	$^{197}_{79}Au$
Copper	29	65	29	36	Cu-65	$^{65}_{29}Cu$
Barium	56	137	56	81	Ba-137	$^{137}_{56}Ba$
Potassium	19	40	19	21	K-40	$^{40}_{19}K$

Chapter 3 Lecture Notes: Compounds

Educational Goals

1. Understand where electrons are located in atoms and how the locations of electrons affect the energy of the atom.
2. Define the term **valence electron** and draw the **electron dot structure** of an atom or ion.
3. Define the term **ion** and explain how the electron dot structure of a s- or p-block element can be used to predict the charge of the monoatomic ion.
4. Given the symbol, be able to name monoatomic **cations** and **anions** (and vice versa).
5. Explain the difference between an **ionic bond** and a **covalent bond**.
6. Understand the structural difference between ionic and covalent compounds.
7. Given the name, be able to write the formulas of ionic compounds and binary covalent compounds (and vice versa).
8. Define the terms **molar mass**, **formula mass,** and **molecular mass** and use these values in unit conversions involving moles and mass.
9. Given the formula, draw the **line bond structures** of diatomic molecules.

The Arrangement of Electrons

Before we learn about compound, we must build on our understanding of atoms and electrons.

Specifically, in the beginning of chapter 3 you will learn:

1) **Where electrons are located in atoms.**
2) **How the location of electrons effect the energy of the atom.**

Scientists used ***light*** to study how electrons are arranged around the nucleus.

Energy, in the form of light or heat, can be ________________by atoms.

Energy is **absorbed** by ________________an electron to a ***new*** area.

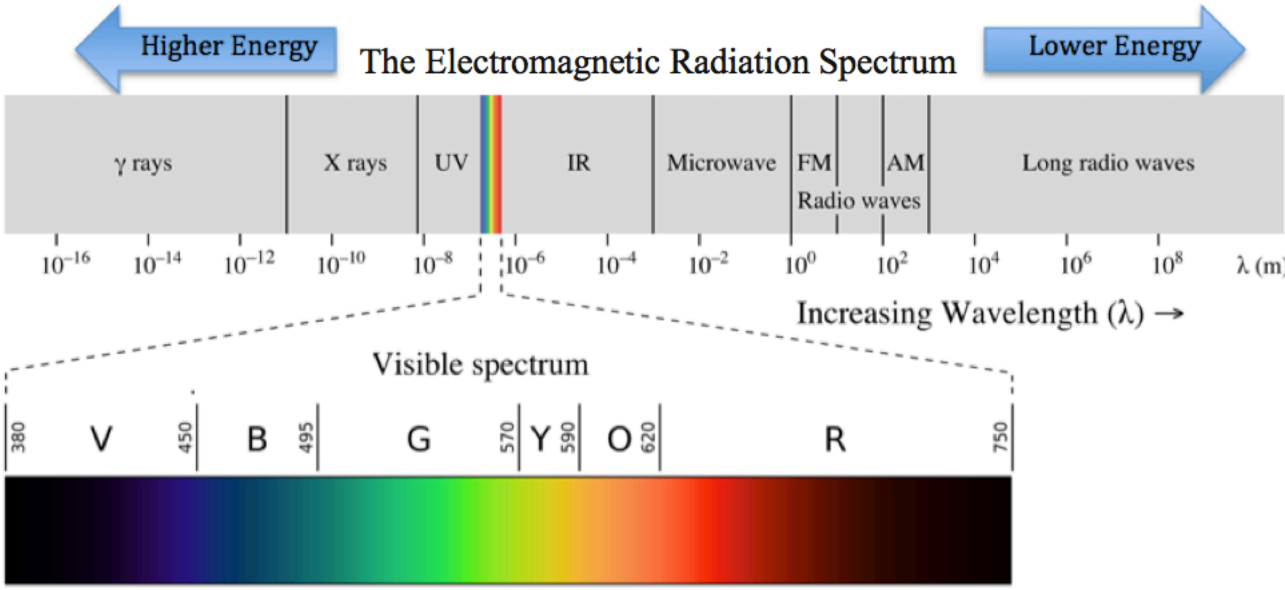
Atoms release energy when electrons move *back* to ________ ____________areas.

- This can happen when an atom collides with another particle.
- Another way this can happen is by an atom emitting ___________.

The Modern Model of the Atom

New scientific laws and models of nature were needed to explain the pattern of light that was emitted by atoms.

Another word for *light* is ***electromagnetic radiation***.



Visible light, the part of the electromagnetic spectrum that can be detected with the human eye, is a small part of the electromagnetic radiation spectrum (see the textbook for colored spectrum).

Short wavelengths correspond to higher energy; longer wavelengths correspond to lower energy light.

If **all energies** of light could be released from excited atoms, then we would expect the pattern of emitted light to look like this:

However, only light with **discrete** (distinct) energies is emitted. For example, the pattern of light emitted from excited hydrogen atoms is:

Our understanding of nature was dramatically changed when Max Planck and Albert Einstein introduced "_______________ _________________."

They proposed that energy is absorbed and emitted by atoms ***only*** in __________ amounts called **quanta**.

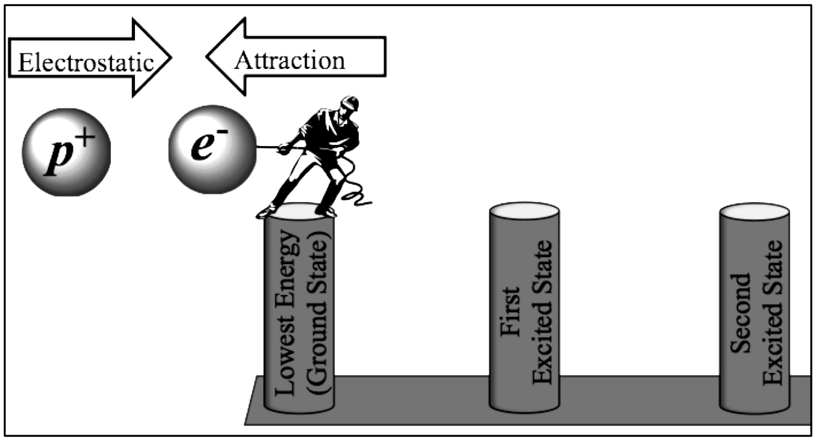
- Another word for "***discrete***" is "***distinct***."

Recall that the light emitted from excited atoms is generated by electrons losing energy as they move from areas further from the nucleus (high energy) to areas nearer the nucleus (low energy). To lose the energy in this process, atoms emit light.

The observation that only discrete energies are emitted from excited atoms is explained using an atomic model that says that the ***electrons can only exist in certain areas and therefore atoms have discrete energies***.

- We say that the energy of atoms is "________________."
- The first scientist to propose a model of the atom where electrons existed in specific regions that had discrete energies was Niels Bohr.

Illustration of the "Quantized" Energy of Hydrogen



When an atom's electron(s) are in the lowest possible energy area, we call this the _________ ________.

- At room temperature, all atoms will exist in their ground state unless *temporarily* excited to a higher energy area by absorbing light.

Absorption of a discrete amount of energy corresponds to the worker **only** being able to move to ***particular areas*** (represented by posts in the illustration above).

When hydrogen's electron is in any other region than the ground state (lowest energy), we call that an _________ ________ of hydrogen.

The excited atom will soon lose energy as the electron moves back to the ground state position. When the energy lost is in the form of light, that light will be the color (wavelength) corresponding to the energy difference between the initial "excited" region and the final, lower energy region.

The Modern Model of the Atom: The Quantum Mechanical Model

You can ***avoid getting lost*** in the detail (and wonder) of nature by focusing on the following two educational goals:

1) **Understand where electrons are _____________ in atoms.**
2) **Understand how the location of electrons affect the ____________ of the atom.**

The Hydrogen Atom

Hydrogen is unique because it has only _________ electron.

Electrons exist in certain **three-dimensional regions** called _____________.

Orbitals can be described by these properties:

1. **The _____________ ____________an electron in a particular orbital is from the nucleus.**
    - Since orbitals are three-dimensional and the electrons move (very quickly) within the orbitals, the distance an electron is from the nucleus is not constant (as it would be in a circular two-dimensional path). Therefore, we talk about the electron's ***average distance*** from the nucleus.

- As hydrogen's orbitals get larger, the average distance of an electron from the nucleus increases, therefore the ____________ ***the orbital*** occupied by an electron, the ____________ ***the energy***. (As described in the illustration at the top of this page).

2. **The three-dimensional ____________ of the orbital.**

- Not only do the sizes of orbitals vary, the shapes of orbitals vary as well.
- When the shapes of orbitals are shown as three-dimensional representations, the shapes represent the region that would contain the electron _________ of the time. The remaining 10% of the time, the electron would be outside of the shape that is shown in the graphic representation.

The Language of Quantum Mechanics

The orbitals are *centered on the* _______________, and are labeled by a ______________.

In a hydrogen atom:

- This number is related to the orbital **size** and the **energy** of an electron in the orbital.
- The orbitals are numbered from **lowest** energy (smallest size) to **higher** energy (larger size).

These numbers are referred to as “energy level,” or “quantum number,” or “quantum level,” or “shell.”

- We will use the term “___________” or “___________ _________” and abbreviate it by using “__.”

In the lowest energy state of a hydrogen atom (the *ground state*), the electron occupies the **n**=1 quantum level.

The **n**=**1** quantum level has _________ orbital.

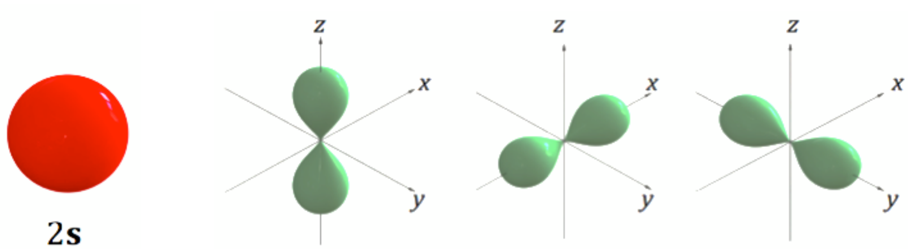
- It is called an _____ orbital.
- **“s”** represents the ***shape* of the orbital**, we use **1s** because **n**=**1**).
- **s** orbitals are _______________ in shape.

An Illustration of a 1s Orbital

- The ____________ is in the **center** of the orbital.

The **n**=**2** quantum level has ________ orbitals.

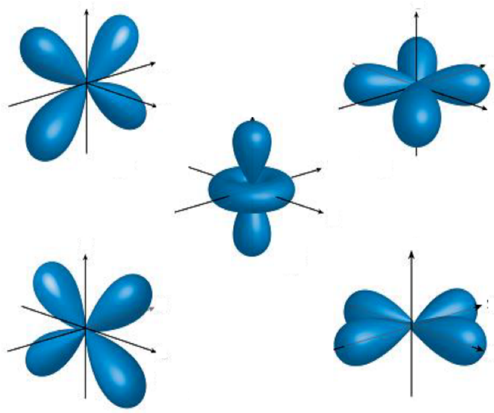
- There is **one 2s** orbital
  - ***All* s** orbitals are spherically shaped.
  - We use **2s** because **n**=**2**.
  - The major difference between the **2s** orbital and the **1s** orbital is that the **2s** orbital is **larger**.
- There are **three 2p** orbitals.
  - **p** represents the shape; we use **2p** because **n**=**2**.
  - The **p** orbitals all have the **same shape** and only differ in *how they are* ______________ around the nucleus.



2s 2p_z 2p_x 2p_y

The **n**=**3** quantum level has __________ orbitals.

- There is **one 3s** orbital, **three 3p** orbitals, and **five 3d** orbitals.
- The shapes of the **3s** and **3p** orbitals are similar to those of the **2s** and **2p** orbitals, respectively, but they are ***larger***.
- The **five 3d** orbital are illustrated below:



As is the case for all orbitals, the **d** orbitals are centered on the nucleus.

The **n**=**4** quantum level has _____________orbitals.

- There is **one 4s** orbital, **three 4p** orbitals, **five 4d** orbitals, and **seven 4f** orbitals.
- The **f** orbitals have shapes that are even more complicated then the **d** orbitals.
- The shapes of the **4s**, **4p**, and **4d** orbitals are similar to those of the **3s, 3p**, and **3d** orbitals, respectively, but they are ***larger***.

The **n**=**5** level has **twenty-five** orbitals.

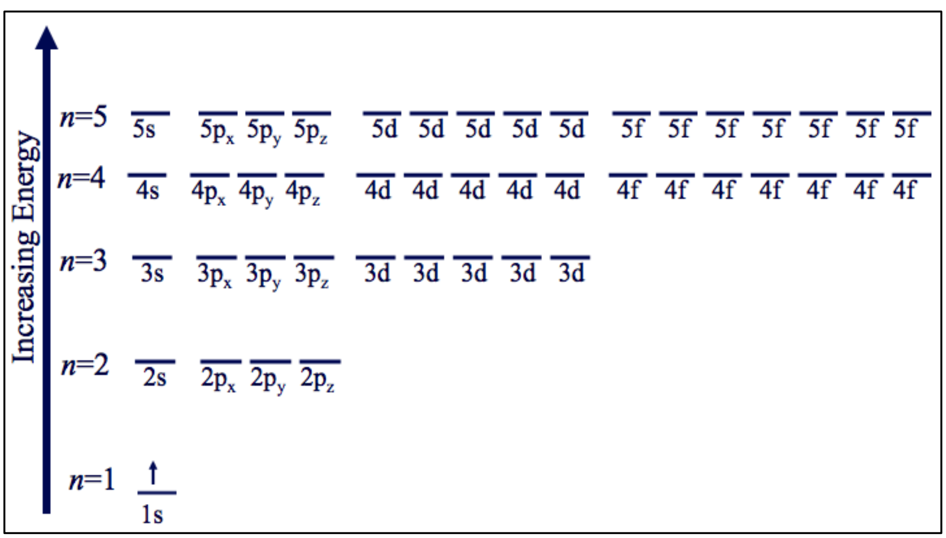
This just keeps going, **n**= **6**, **7**, **8**, etc.

Although quantum levels with **n** > **4** contain orbitals other than **s**, **p**, **d**, and **f**, these other orbitals are n*ever occupied by electrons of any element in its ground state*.

- The only time an electron can occupy any of those orbitals will be if the atom absorbs energy.

Energy Level Diagram for Hydrogen

We can organize the various atomic orbitals according to their energy in an illustration called an **energy level diagram**. The energy level diagram for the first five quantum levels (**n** = 1-5) of a **hydrogen atom** is shown below.

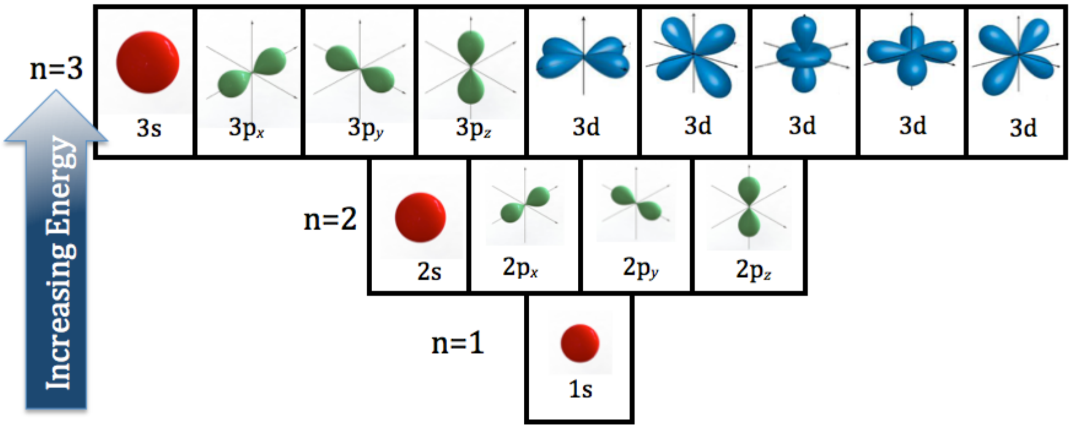


In an **energy level diagram,** we draw a short horizontal line that is labeled for each __________.

The orbitals are arranged, from bottom to top, in order of increasing ___________.

An electron is depicted as an ___________ above the line that represents the orbital occupied by it.

Let’s compare the *energy level diagram* to a *skyscraper*, we will call this our **skyscraper model**.



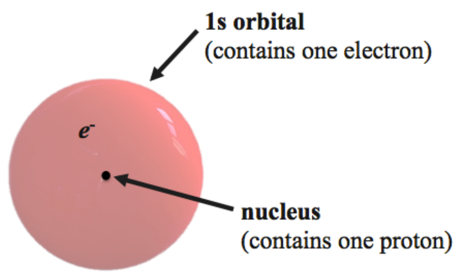
Skyscraper photo source: Wikimedia Commons, Author: Avala, CC-BY-SA, http://creativecommons.org/licenses/by-sa/3.0/legalcode

The different ____________ of the skyscraper represent the quantum levels (**n**).

- The first floor is the lowest energy floor so it would correlate with **n**=1, the ground state.

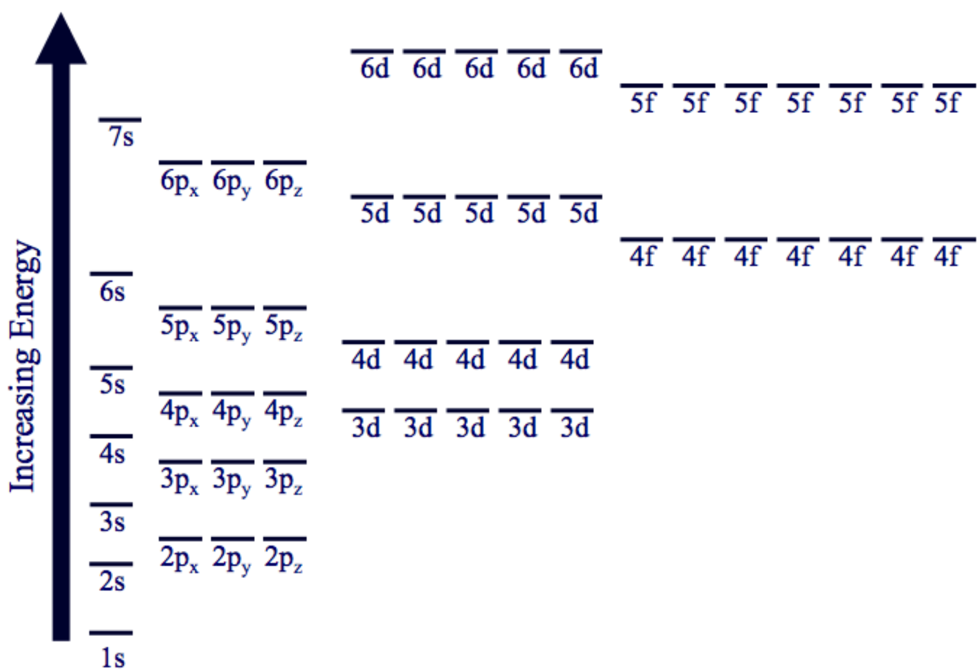
___________ on a particular floor are analogous to the various orbitals in a particular *quantum level*.

Model of a Hydrogen Atom

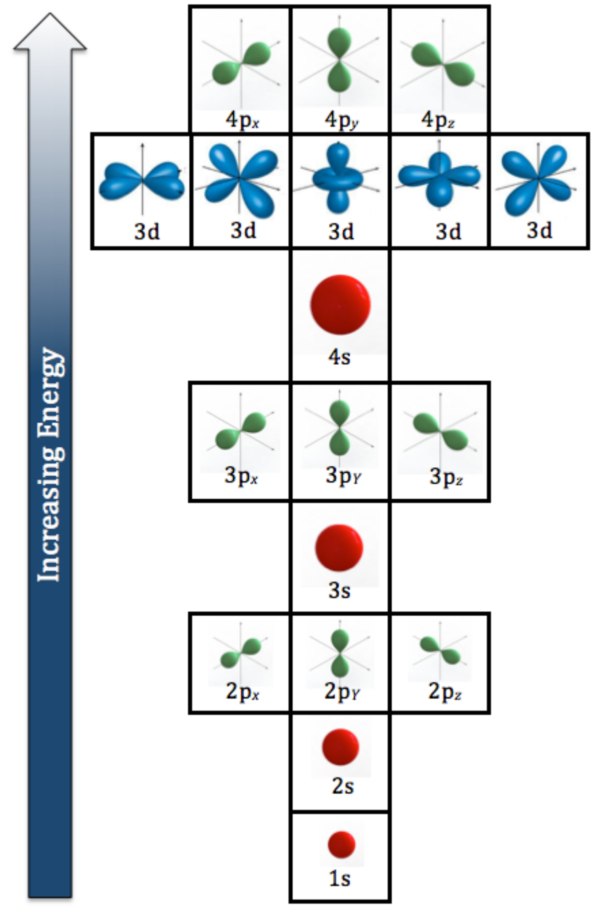


Atomic Model for Multi-Electron Atoms

Energy Level Diagram for Multi-Electron Atoms



Skyscraper Model for Multi-Electron Atoms



We live in a universe where matter tends to exist in the *lowest possible energy state*.

An informal way to state this is:

Nature wants everything to be at the ___________ possible energy.

Electrons are arranged (configured) into the orbitals of multi-electron atoms in the way that results in the **lowest** possible energy.

Nature does this by obeying the following principles:

1) **The Aufbau Principle**

"Aufbau" (German) means *build-up* or *construct*.
The *aufbau principle* states that an electron occupies the ***lowest energy orbital that can receive it***.

2) **The Pauli Exclusion Principle**

An orbital can hold a maximum of ________electrons.

Electrons have a quantum mechanical property called **spin**.

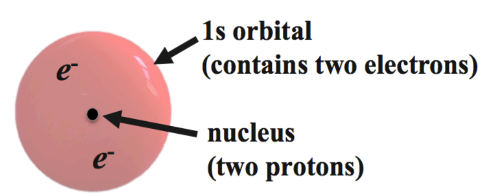
We call the spin states "**up**" or "**down**."

- *When two electrons occupy the same orbital, one electron has spin "**up**" the other has spin "**down**."*

This is all that you need to know about spin to understand all of the concepts covered in this textbook. You may find it interesting that spin is responsible for magnetic properties of matter. In fact, spin is the reason that electrons behave as tiny magnets!

Example:
Energy Level Diagram for a Helium Atom

Model of a Helium Atom



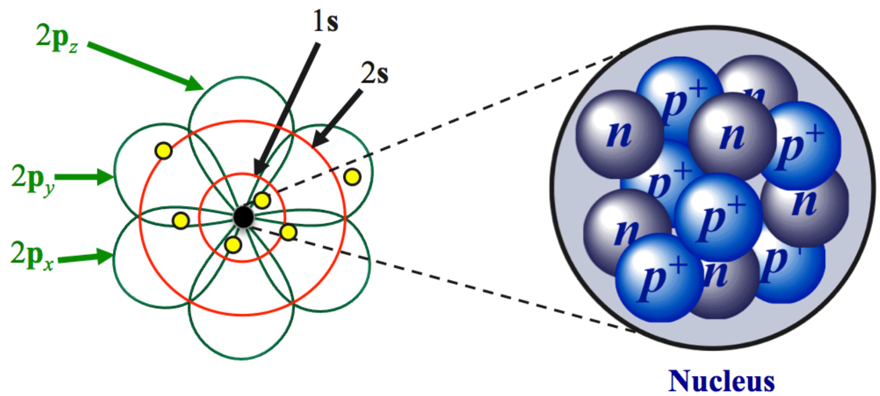
3) Hund's Rules

When electrons are configured into orbitals that all have the *same energy,* a ***single electron*** is placed into **each** of the equal-energy orbitals before a second electron is added to an occupied orbital.

When electrons are configured into a *set of orbitals* that all have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all "up").

Example: Energy Level Diagram for a Carbon Atom

Drawing of a Carbon-12 Atom



Understanding Check:
Energy Level Diagram for a Neon (Ne) Atom

Understanding Check:
Energy Level Diagram for an Iodine (I) Atom

Valence Electrons

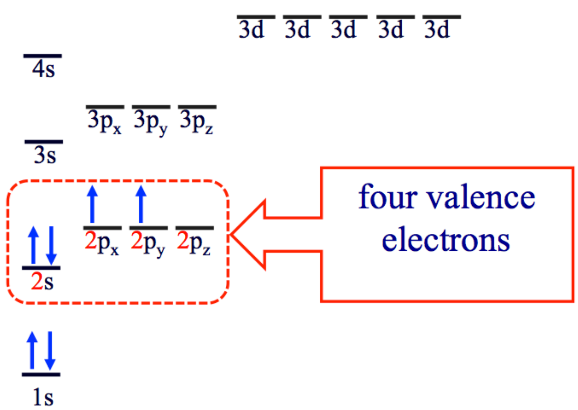
Valence electrons are the electrons held in the ___________________________ ***shell*** (largest "**n**").

Language Reminder: "***shell***" = "***quantum level***" = "***energy level***"

– Valence electrons are furthest away from the _____________________.

- It is important to know how many valence electrons are in an atom because:
  – These are the electrons that are involved in ______________________ ____________________ to other elements to form ______________________.
  – These are the electrons that elements lose to become ________________.

Example: How many valence electrons do **carbon** (C) atoms have?



Understanding Check: How many valence electrons do **oxygen** (O) atoms have?

Short-Cut for Determining the Number of Valence Electrons

Elements are arranged in the periodic table according to the number of valence electrons.

For **s-** and **p-block** elements, all elements in the same periodic ____________ (group) have the *same number of valence electrons* as all others in that column.

The group numbers for the columns *represent* the number of valence electrons contained in those atoms.

I																	VIII
1 H	II											III	IV	V	VI	VII	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg						Transition Metals					13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt									

(Inner) Transition Metals

Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Different elements with the same number of valence electrons are said to be ________________.

Example of isoelectric elements: oxygen and sulfur

Isoelectric atoms often behave in similar ways. For example, oxygen atoms often chemically "bond" to two hydrogen atoms to form water (H_2O); sulfur atoms, also often "bond" with two hydrogen atoms to form hydrogen sulfide (H_2S).

Understanding Check: Use the periodic table to determine the number of valence electrons in each of these types of atoms:

a. hydrogen (H)

b. nitrogen (N)

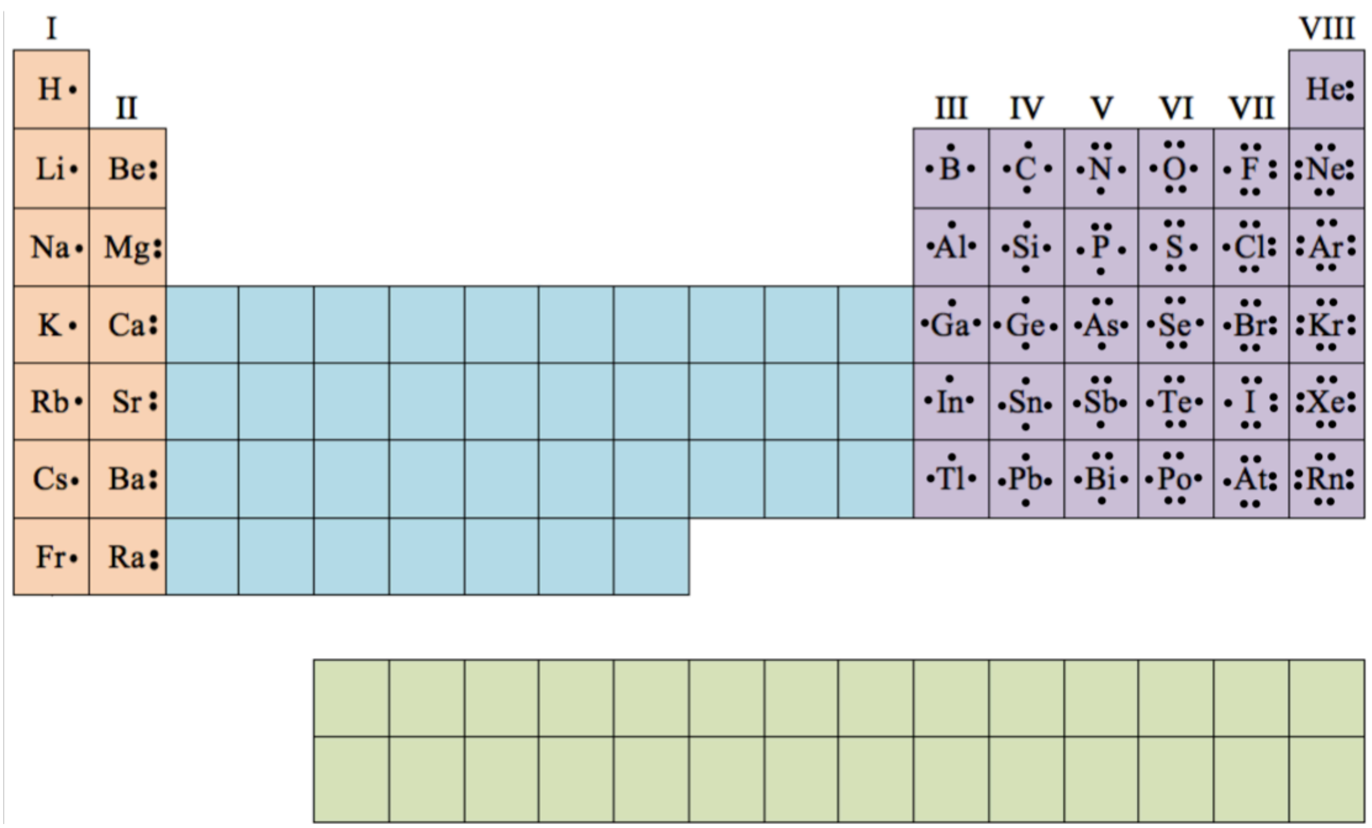
c. bromine (Br)

d. krypton (Kr)

Electron Dot Structures

Electron dot structures show the number of *valence electrons* that an atom carries.

- In these structures, *valence electrons are represented by* _____ drawn next to an element's symbol.

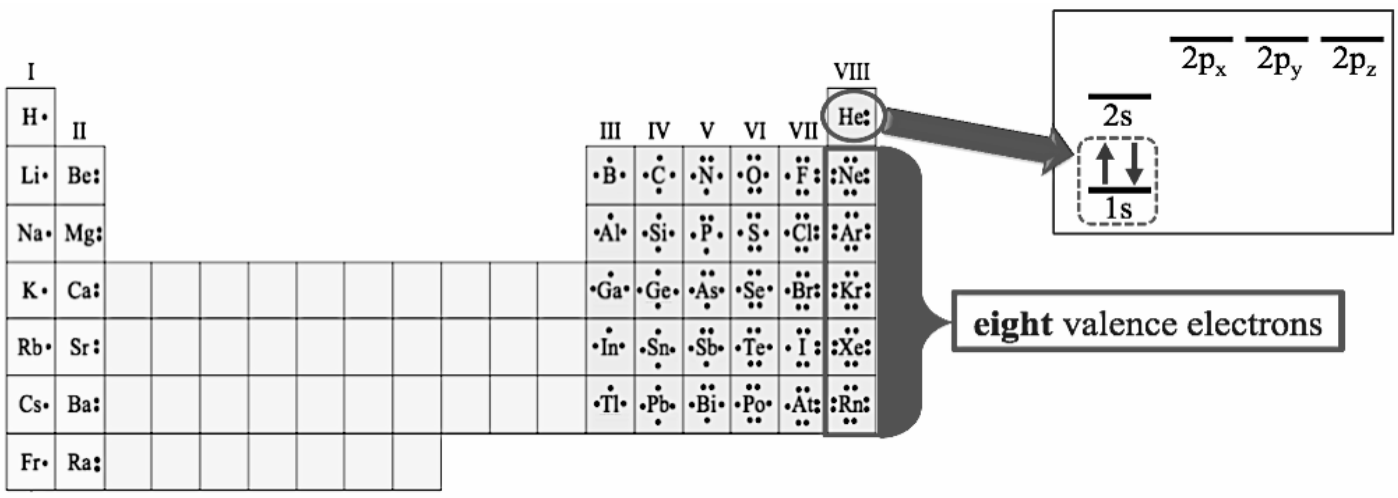


Noble Gases and the Octet Rule

The group VIII elements (He, Ne, Ar, Kr, Xe, and Rn) are called _________ ***gases.***

He, Ne, Ar, Kr, Xe, and Rn belong to the **noble gas family**, which gets it's name from the fact that these elements are resistant to change and, with few exceptions, do not lose or gain electrons.

The resistance to change (stability) of the noble gases is related to having their outermost quantum level (***shell***) completely _______________with electrons



Helium's outermost shell (the n=1 quantum level) is completely filled with its ***two*** electrons.

All of the other noble gas elements have completely filled outermost shells with _________ electrons.

This stability of the noble gas elements that have ***eight electrons*** in their outermost shell led to what chemists call the _________ ________.

Most substances around us do not exist as individual atoms. Atoms will "bond" with other atoms to form compounds such as water (H_2O), carbon dioxide (CO_2), and table salt (sodium chloride). In the remainder of chapter 3, we will discuss the nature of this "bonding" of atoms to other atoms.

The Octet Rule is quite useful in predicting and understanding bonding patterns in chemical compounds.

The Octet Rule

Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an ***octet*** (eight) of electrons in its outermost shell.

There are exceptions to the octet rule. An important exception that we will always use is for ***hydrogen*** and ***helium***.

- Hydrogen and helium have filled outer shells (are stable) with just ***two*** electrons because their outermost level (**n**=1) has only one orbital.

Ions

Atoms have the same number of electrons as protons and are therefore ***electrically neutral***.

An _________ is a small particle that has an ***electrical charge***.

Atoms can ***gain or lose*** _____________ to become **ions**.

Metal atoms can _________ electrons to form ***positive ions***.

If an atom *loses* one or more electrons, it will then have more protons than electrons and have an overall *positive charge*.

- *Positive ions* are called ___________.

Nonmetal atoms can _______ electrons to form ***negative ions***.

If an atom *gains* one or more electrons, it will then have more electrons than protons and have an overall *negative charge*.

- *Negative ions* are called _____________.

The Octet Rule in the Formation of Ions

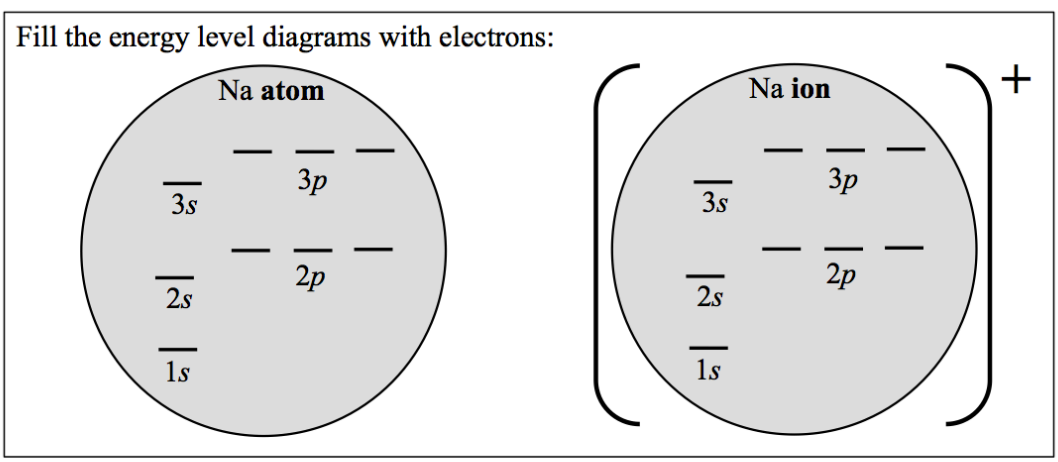
The ***Octet Rule*** can be used to predict the formation of ions.

Very often, ions are formed such that the ***ion*** has an ___________ in its outermost shell.

This tendency will allow us to predict the ___________ of the ion that is formed for particular elements.

Example: Let's do a Cation - Sodium (Na)

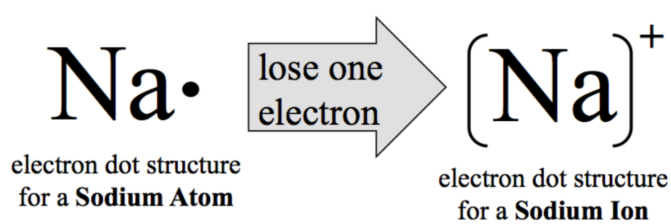
- A sodium atom has ______ protons and _____ electrons.
- How many valence electrons does the sodium atom have? _____
- How many valence electrons does sodium want? _____



When sodium loses an electron, it has an octet of electrons in its outer shell.

Sodium will lose________ electron to become a sodium ion (Na^+).

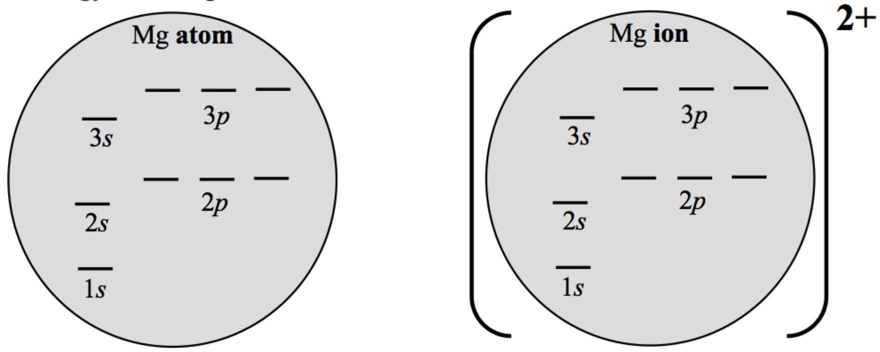
- Sodium has one valence electron
- There are two ways to have a filled octet:
  - 1) Add 7 electrons
  - 2) Remove one electron
- It is easier to remove one electron!



Example: Let's do Another Cation - Magnesium (Mg)

- A magnesium atom has ______ protons and _____ electrons.
- How many valence electrons does the magnesium atom have? _____
- How many valence electrons does magnesium "want?" _____

Fill the energy level diagrams with electrons:



When magnesium loses two electrons, it has an octet of electrons in its outer shell.

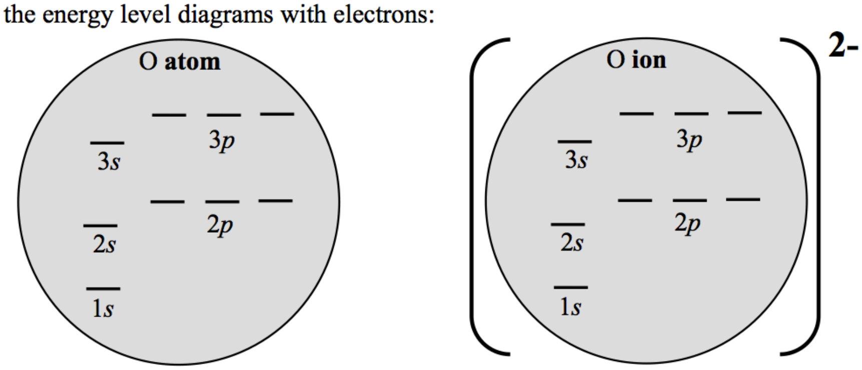
Magnesium will lose ______ electrons to become a magnesium ion (Mg^{2+}).

Understanding Check: Based on the octet rule, what would be the charge of an aluminum ion?
HINT: Begin with the energy level diagram (or the number of valence electrons) for an aluminum atom.

Example: Let's do an Anion - Oxygen (O)

- An oxygen atom has ______ protons and ______ electrons.
- How many valence electrons does the oxygen atom have? ____
- How many valence electrons does oxygen want? _____

Fill the energy level diagrams with electrons:

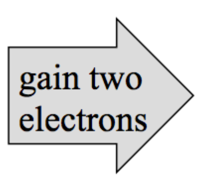


When oxygen gains two electrons, it has an octet of electrons in its outer shell.

Oxygen will gain_______ electrons to become an oxide ion (0^{2-}).

The ***electron dot structure*** can give us the same conclusion!

Draw an electron dot structure for an **Oxygen Atom**:

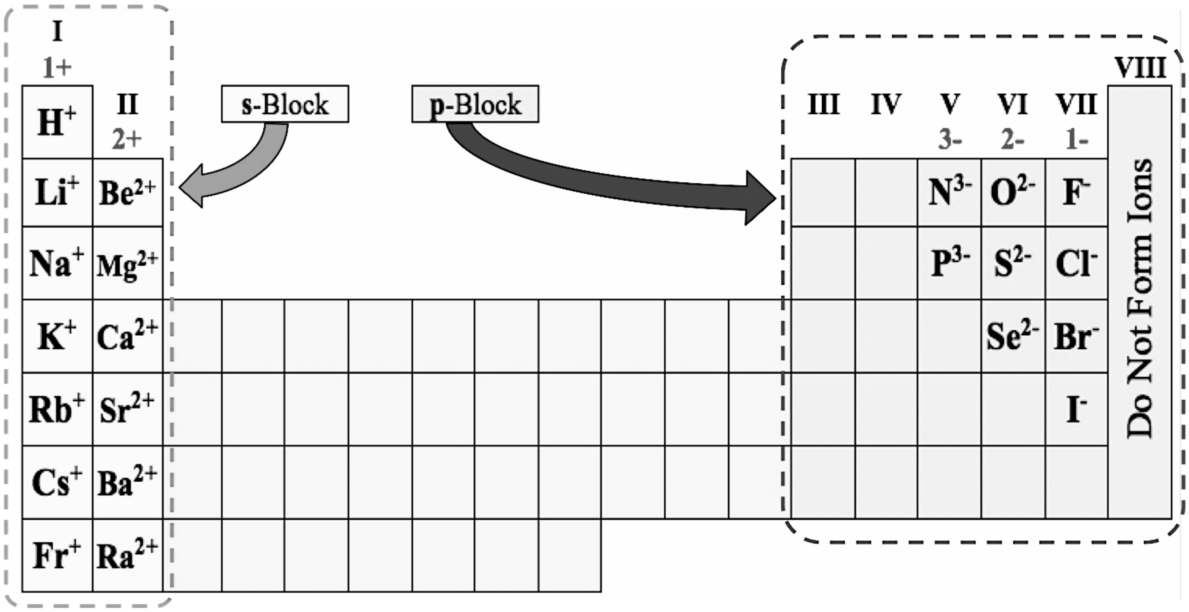


Draw an electron dot structure for an **Oxide Ion**:

Oxygen has 6 valence electrons, if we add two electrons; its outer shell will have an octet.

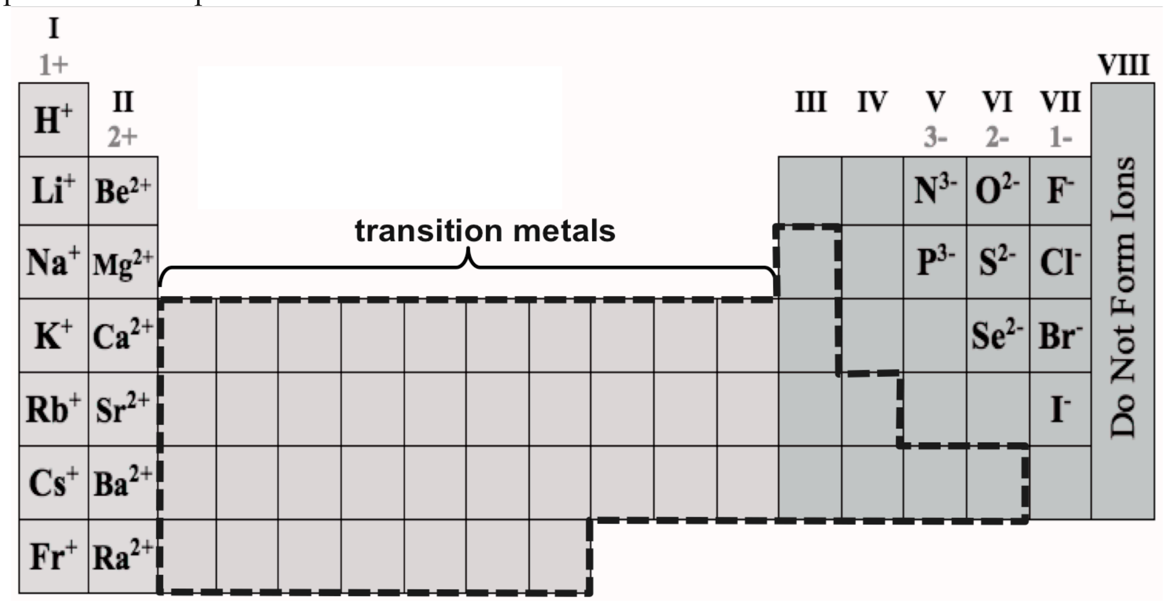
Understanding Check: What would be the charge of an **ion** formed from a **chlorine** atom?
HINT: Begin with the electron dot structure for a chlorine atom.

We can determine the charge of an ion formed from ***s-block elements*** and ***p-block nonmetals*** from the number of valence electrons in those elements, and therefore by their ***location*** on the periodic table.



Periodic Group	Number of Valence Electrons of the Element	Number of Electrons Gained or Lost in Ion Formation	Charge of Ion Formed
s-Block Elements			
Group I	1	Lose 1 electron	1+
Group II	2	Lose 2 electrons	2+
p-Block Nonmetal Elements			
Group III	There are no Group III nonmetals (only metals and metalloids)		
Group IV	4	Do not form ions, high energy to gain **or** lose 4 electrons!	
Group V	5	Gain 3 electrons	3-
Group VI	6	Gain 2 electrons	2-
Group VII	7	Gain 1 electron	1-
Group VIII	8	Do not form ions, noble gas atoms have filled outer shells.	

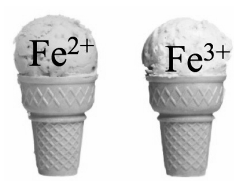
The charge of the ions formed from the **transition metals** and **p-block metals** ***cannot*** always be predicted by their position in the periodic table.



Many of these elements can form ***more than one*** type (charge) of ion.

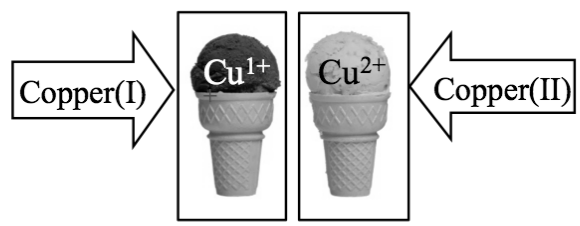
Example: Iron (Fe):

Iron (Fe) ions can come as Fe^{2+} or Fe^{3+}



Example: Copper (Cu):

Copper (Cu) ions can come as Cu^{1+}or Cu^{2+}



To differentiate the various charge states of ions when reading or writing their names, we use ___________ numerals corresponding to the **charge** after the element name.

- When saying the ion's name, one would say "copper one" for Cu^{1+} and "copper two" for Cu^{2+}.

We ***only*** use the Roman numeral for ions that can ***exist in more than one charge state***.

Some of the **transition metals** and **p-block metals** only exist in **one** charge state.

- For example, **cadmium ions** only exist as $\mathbf{Cd^{2+}}$.

Roman numerals ***are not used*** when the metal cations *have just one charge state.*

Since the charges of many of the transition metal and p-block metal ions ***cannot*** be easily predicted from their positions on the periodic table, and many can have more than one charge, we must refer to tabulated list for the charges.

The table below lists the charges for some transition metals and p-block ions. **You do not need to memorize the metal names and charges in this table**; I will give you this table for with your exams.

Charges for Some Transition Metal and p-Block Metal Ions

Ions that occur with only one charge			
Name	Charge	Name	Charge
aluminum ion	Al^{3+}	cadmium ion	Cd^{2+}
silver ion	Ag^{+}	zinc ion	Zn^{2+}
Ions that occur with multiple charges			
Name	Charge	Name	Charge
copper(I) ion	Cu^{+}	tin(II) ion	Sn^{2+}
copper(II) ion	Cu^{2+}	tin(IV) ion	Sn^{4+}
iron(II) ion	Fe^{2+}	lead(II) ion	Pb^{2+}
iron(III) ion	Fe^{3+}	lead(IV) ion	Pb^{4+}
cobalt(II) ion	Co^{2+}	mercury(I) ion	Hg^{+}
cobalt(III) ion	Co^{3+}	mercury(II) ion	Hg^{2+}

This table does not contain data for the all ions formed by ***all*** of the transition and **p**-block metal cations, however it includes the ions that you will need in order to solve and understand any of the examples and problems in this course.

Naming Monatomic Ions

A ***monatomic ion*** is an ion that is made when a ___________ **atom** gains or loses electron(s).

Naming Monatomic *Cations*

Cations use the name of the element, followed by the word "ion."

- **Examples:**
  - **Na^{+}** is referred to as a sodium ion.
  - **Mg^{2+}** is referred to as a magnesium ion.

For monatomic cations that can occur *with multiple charges*, indicate the charge using Roman numerals *after* the element's name.

- **Examples:**
  - **Fe^{2+}** is referred to as an iron(II) ion
  - **Fe^{3+}** is referred to as an iron(III) ion

Naming Monatomic *Anions*

Anions are named by changing the *suffix* (ending) of the name to "-________."

- **Examples:**
  - **F^{-}** is referred to as a fluor**ide** ion.
  - **O^{2-}** is referred to as an ox**ide** ion.

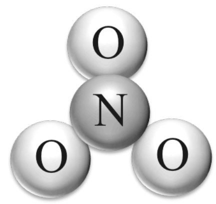
Polyatomic Ions

Several atoms often "*stick*" (bond) together to form a small particle.

If the resulting particle has the s*ame number of protons as electrons*, then it will be ***electrically neutral***, and we call the particle a ___________________.

If, on the other hand, there is an *excess of protons or an excess of electrons* in the particle, then it will have an ***overall electrical charge***, and we call the particle a ___________________ **ion**.

Example of a Polyatomic Ion: Nitrate Ion



Nitrogen (7 electrons, 7 protons)
Oxygen (8 electrons, 8 protons)
Oxygen (8 electrons, 8 protons)
Oxygen (8 electrons, 8 protons)
\+ **one extra electron**

Nitrate Ion
NO_3^-

The table below lists the names and charges for some polyatomic ions. **You do not need to memorize this table**; I will give you this table for with your exams.

Some Polyatomic Ion Names and Charges

POLYATOMIC CATIONS	
H_3O^+ hydronium ion	NH_4^+ ammonium ion
POLYATOMIC ANIONS	
OH^- hydroxide ion	HSO_4^- hydrogen sulfate (or bisulfate) ion
CO_3^{2-} carbonate ion	PO_4^{3-} phosphate ion
HCO_3^- bicarbonate (also called hydrogen carbonate) ion	HPO_4^{2-} hydrogen phosphate ion
NO_2^- nitrite ion	$H_2PO_4^-$ dihydrogen phosphate ion
NO_3^- nitrate ion	CrO_4^{2-} chromate ion
SO_3^{2-} sulfite ion	$Cr_2O_7^{2-}$ dichromate ion
SO_4^{2-} sulfate ion	$C_2H_3O_2^-$ acetate ion (sometimes written as $CH_3CO_2^-$)
	CN^- cyanide ion

Chemical Compounds

Compounds: ____

____

Each compound has the same ________________________of the same elements.

- Example: Water = 2 hydrogen atoms and 1 oxygen atom (Ratio H:O = 2:1)

Chemical bonds

Atoms can ***bond*** with other atoms, and ions can ***bond*** with other ions to form ***compounds*** such as water (H_2O), carbon dioxide (CO_2), and table salt (sodium chloride).

Chemical bonds are the ***electrical attractive*** _______________that hold atoms or ions together in a compound.

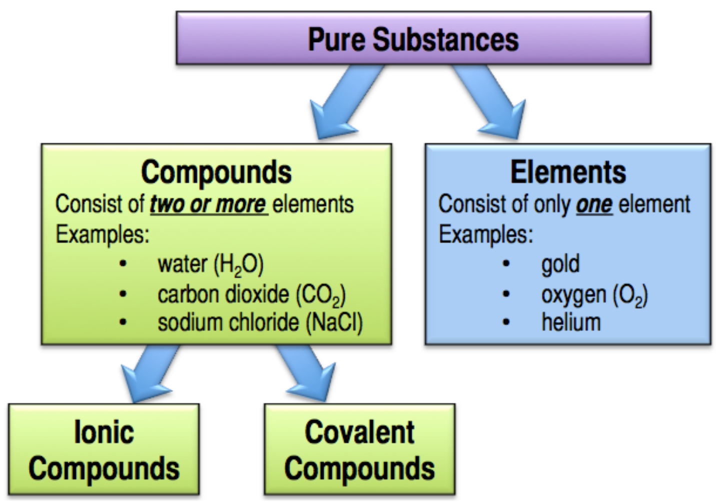
There are *three types* of **chemical bonding**:

1) **Covalent Bonding**
2) **Ionic Bonding**
3) **Metallic Bonding**

In this chapter, you will learn about the first two types, *covalent bonding* and *ionic bonding*. You will learn about *metallic bonding* in chapter 5.

Some Terminology

All matter can be classified as either ***mixtures*** or ***pure substances***. You will learn about mixtures in chapter 6. ***Pure substances*** are described in the chart *on the right*:



Chemistry is the study of matter and the ***changes*** it undergoes.

______________ ***changes***, such as *melting* or *boiling*, result in changes in *physical properties* and ***do not*** involve the formation of new *pure substances*.

- For example, the melting of ice is simply H_2O being changed from the *solid* phase to the *liquid* phase. The chemical bonds between oxygen and hydrogen atoms do not change in that process.

_____________ ***changes***, on the other hand, result in the formation of ***new*** *pure substances*.

- To make a **new** pure substance, ***chemical bonds must be*** ____________ ***and/or new chemical bonds are*** ____________.
- This happens in a process called a **chemical reaction**, which we will study in chapter 6.

A *major principle* of chemistry is that the observed (macroscopic) properties of a substance are related to its "microscopic" structure.

- The *microscopic* structure entails details such as the kind of atoms/ions and the pattern in which they are ***bonded*** to each other.

Covalent Chemical Bonding

Covalent bonding is defined as the chemical bonding *force* that results from the ____ between two atoms.

The resulting collection of atoms results in the formation of either ________________ or ***polyatomic ions***.

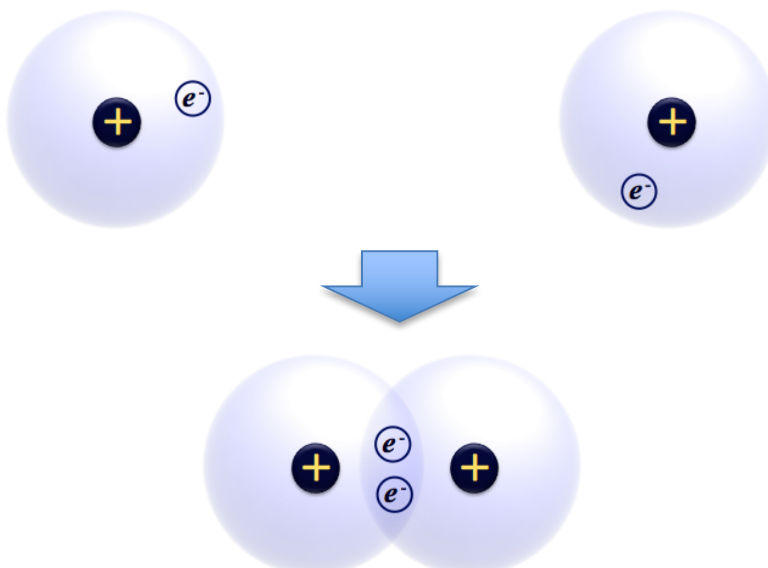
A **molecule** is an electrically ______________ group of atoms ***held together by covalent bonds***.

Covalent bonding occurs between _________________atoms.

Covalent bonding occurs because the bound atoms are at a _________ energy than the unbound atoms.

Why does sharing of electron pairs result in an attractive electrostatic force capable of holding atoms together?

Consider the two ***hydrogen atoms*** coming together to form a covalent bond.



In covalent bonding, the atoms __________ *electron pairs*.

The shared electron pair spends significantly ***more*** _______ in the area between the positive nuclei of the hydrogen atoms than in other regions.

The electron pair between the nuclei create a **positive-negative-positive** electrostatic attractive "sandwich" and this ___________ holds the atoms together.

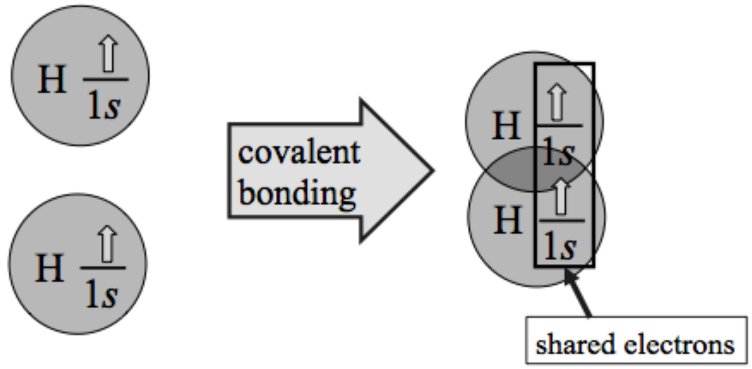
The Octet Rule in the Formation of Molecules

The positive-negative-positive model **cannot** explain why a covalent bond ***does not*** form between two *helium* atoms.

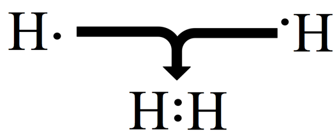
The octet rule in the formation of molecules is: ***molecules tend to form such that the atoms are surrounded by an octet (eight) of valence electrons (except for hydrogen and helium that have two electrons).***

Example: H_2
(recall that H and He are stable with two valence electrons)

When a covalent bond forms, each hydrogen atom "feels" **two** electrons in its outermost shell.



The H_2 covalent bond can also be illustrated with electron dot structures.



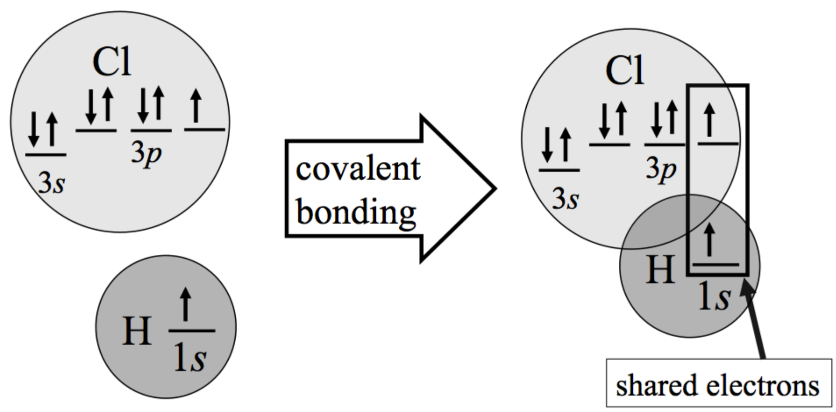
The two electrons _____________the atoms are shared in a covalent bond.
Chemist use a line to represent ____ electrons in a covalent bond.

$$\mathrm{H{-}H}$$

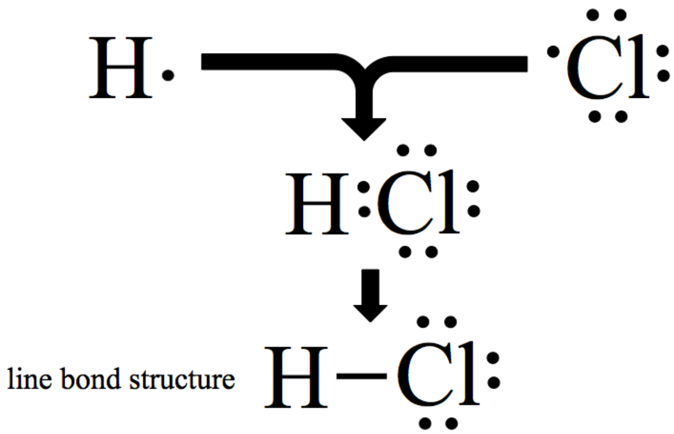
These drawings are called __________________________________.

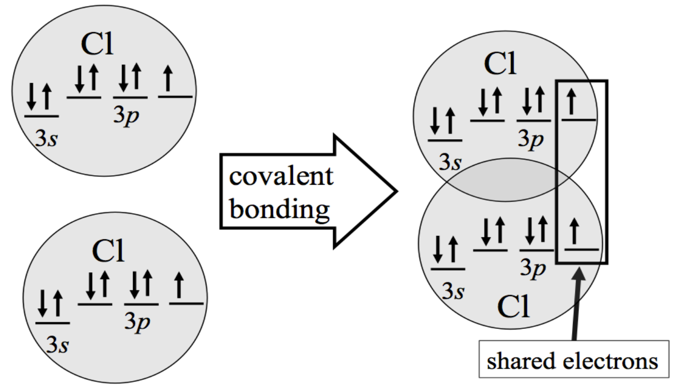
Let's do another example: **Hydrogen Chloride (HCl)**

When a covalent bond forms, the hydrogen atom "feels" **two** electrons in its outermost shell, and the chlorine atom "feels" **eight** electrons in its outermost shell.



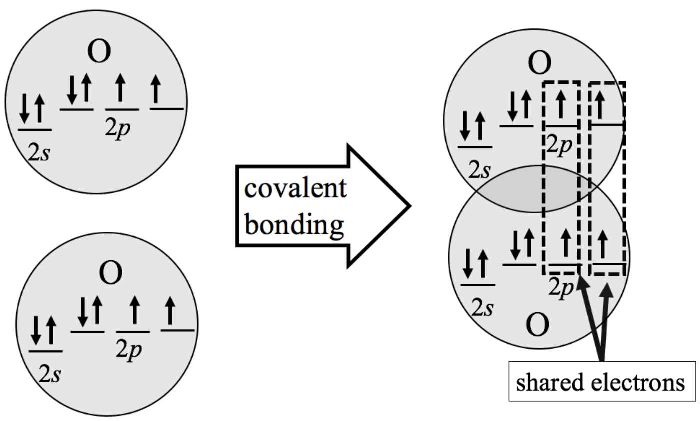
The HCl covalent bond can also be illustrated using electron dot structures.





Let’s do another example: **Cl_2 (chlorine gas)**.

When a covalent bond forms, each chlorine atom “feels” **eight** electrons in its outermost shell.



You try it:
Draw the ***line bond structure*** for Cl_2.
- Start with the electron dot structure for two Cl atoms.

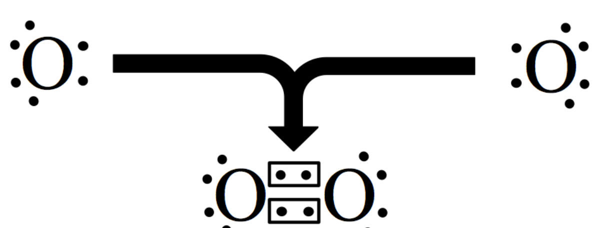
Let’s do **oxygen gas (O_2)**.

In O_2, ***two pairs*** of electrons are shared.

When a covalent bond forms, each oxygen atom “feels” **eight** electrons in its outermost shell.

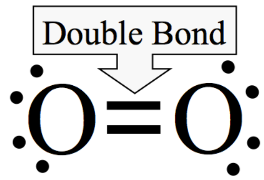
Let’s draw the line bond structure for **oxygen gas (O_2)**.

- Oxygen atoms have 6 valence electrons.
- We will rotate the electrons so they can form bonding pairs.



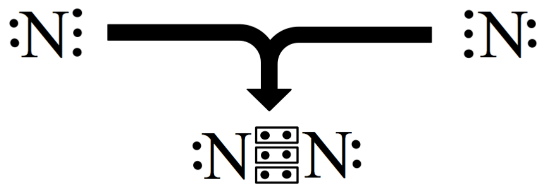
We use lines to represent _____________electron pairs.

When atoms are bonded with 2 pairs of electrons it is called a ____________ __________.



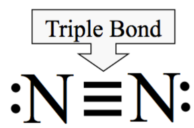
Let's draw the line bond structure for nitrogen gas (**N_2**)

- Nitrogen atoms have 5 valence electrons.
- We will rotate the electrons so they can form bonding pairs.



We use lines to represent electron pairs.

When atoms are bonded with **3 pairs** of electrons it is called a ___________ _________.



Naming Covalent Compounds

The covalent bonding that we will see in this course will *always* involve _______________ ***elements* only**.

The nonmetal atoms can share electrons to form **molecules** (***molecular compounds***) or **polyatomic ions**.

A chemical substance whose simplest units are **molecules** is called a ___________________ ***compound***.

When discussing molecules we use a ________________ ______________ that shows the *types (elements)* and *numbers of atoms* that make up a single molecule.

The number of atoms of each element contained in the molecule is written as a ***subscript*** after the element's symbol.

- **Examples:**

line bond structure	molecular formula
H–H	H_2
$H-\ddot{\underset{\cdot\cdot}{O}}-H$	H_2O

When there is only **one atom** of a particular element present in a molecule the subscripted "**1**" is ***omitted*** for that element.

Some molecules ***only*** contain ***one*** element, for example **H_2**, **Cl_2**, and **O_2**.

- These molecules often take the name of the elements they contain.
- **Examples:**

molecular formula	name
H_2	hydrogen
O_2	oxygen

Naming Binary Covalent (Molecular) Compounds

____________ **covalent compounds** contain only ***two*** ______________ (the "**bi-**" prefix indicates "***two***").

- Examples of binary covalent compounds are **HCl**, **H_2O**, and **CO_2**.

Educational Goals:

Given the **name** of a ***binary covalent molecule***, be able to write the **molecular formula**.

Given the **molecular formula** of a ***binary covalent molecule***, be able to write the **name** of the molecule.

Method for *Naming* Binary Covalent (Molecular) Compounds

1. List the name of the first element in the formula.
2. List the second element and add the –**ide** ____________.
3. Use Greek _________ to indicate the number of each atom in the formula.
   - **Exception:** If there is just one atom of the _________ element in the formula, do not use **mono-** for the ***first element in the name***.
     - **Example: CO_2**
       ~~mono~~carbon dioxide → carbon dioxide
   - The "**o**" or "**a**" at the ***end of the Greek prefix*** is omitted when the element's name begins with a vowel.
     - **Example: CO**
       carbon mon~~o~~oxide → carbon monoxide

Greek Prefix	Number
mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7
octa	8
nona	9
deca	10

Example Problem:

Name the following compound: **CCl_4**

Answer: carbon tetrachloride

Understanding Check

Write the **names** of the following molecules:

CF_4 ________________________________

N_2O ________________________________

SF_6 ________________________________

Method for Writing the *Molecular Formula* of a Binary Covalent Compound

1. Write the symbol of the *first element* in the compound's name, then the symbol of the *second element* in the compound's name.
2. Indicate *how many atoms* of each element the molecule contains using ***subscripts*** after the atomic symbol.
   - The *numbers of atoms* are given in the Greek prefixes in the molecule's name.
   - NOTE: If there is no Greek prefix in front of the first element in the name, that means the number is **1**.

Example:
Write the *molecular formula* for **dinitrogen tetrafluoride**.

N_2F_4

Understanding Check

Write the **molecular formula** for the covalent compounds:

- nitrogen trichloride ______________
- dinitrogen pentoxide ______________
- sulfur dioxide ____________

For covalent compounds with ***more than two types of atoms***, we use *common names* or IUPAC system names. You are not responsible for knowing *common names*. You will learn some IUPAC system names in later chapters.

Examples of *common names*:

- Glucose ($C_6H_{12}O_6$)
- Acetone (C_3H_6O)

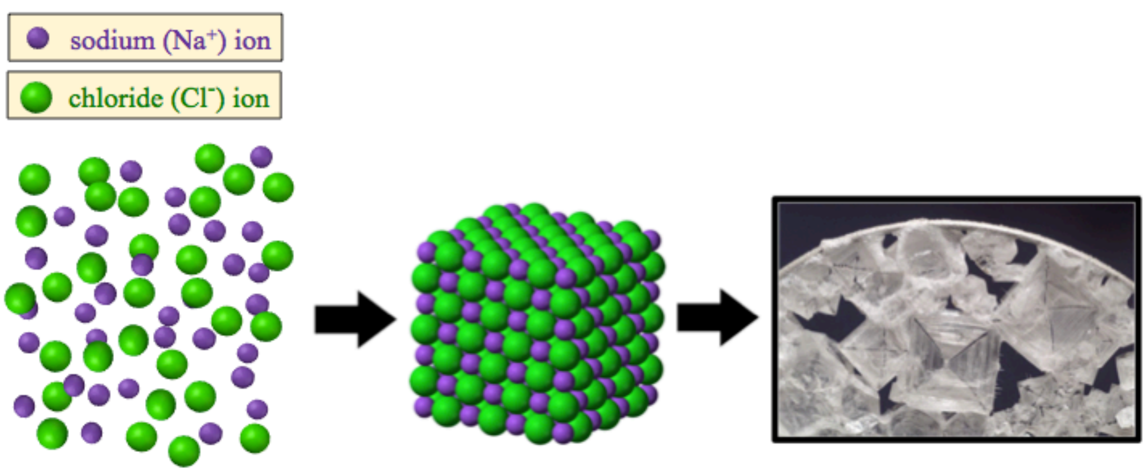
Ionic Compounds

Definition of **ionic bonding**: Chemical bonding that results from the electrostatic attraction between _____________ **numbers** of ***cations*** and ***anions***.

- Compounds composed of ions are called **ionic compounds**.

Example of an ionic compound: sodium chloride (NaCl)

Many sodium ions combine with **many** chloride ions in a *three-dimensional pattern* that minimizes the distance between the oppositely charged cations and anions and maximizes the distance between the like-charged particles.



We call this structure a ______________ or ***crystal*** _____________.

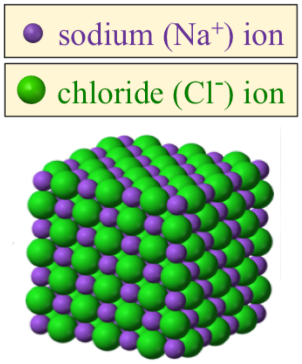
It is this regular, repeating structure on the scale of the individual ions that give crystals the interesting geometrical shapes that we see on the macro-scale when we look at them with our eyes or with a microscope.

Ionic bonding (ionic compounds) results from:

- Combining **metal ions** with **nonmetal** ions.
- Combining **polyatomic ions** with ***other*** **ions**.

The cations and anions will combine in a ratio such that the ***total*** of the *positive* (+) and *negative* (–) *charges* equals _____________!

- **Example:** Sodium Chloride (NaCl)



Sodium ions have a charge of 1+

Chloride ions have a charge of 1-

They combine in a 1-to-1 __________ in the crystal

For every sodium ion, there is one chloride ion!

The charges _________ up to ZERO!

Formula Units

The use of *molecular formulas* would not make sense for ionic compounds; they do not form molecules, instead they form crystals.

We write ____________ __________ **(**as opposed to ***molecular formulas*****)** for **ionic compounds**.

The *formula unit* looks like the molecular formula used for covalent compounds, however it means something ***entirely*** different.

The *formula unit* uses ***subscripted numbers*** after the ion's symbol that indicate the ***ratio*** that the cations and anions combine in the ionic crystal.

- As in the case of molecular formula, when a subscript would have a value of "1," the subscript is omitted.
- We write the cation symbol first followed by a numerical subscript (if needed), then we write the anion symbol followed by a numerical subscript (if needed).

Example: For sodium chloride, since sodium ions and chloride ions combine in a **one-to-one ratio**, we write the formula unit of sodium chloride as:

$$NaCl$$

Example: *Calcium* ions combine with *fluoride* ions to form an **ionic compound**.

Calcium ions have a charge of 2+

Fluoride ions have a charge of 1-

They combine in a ___________ ratio in the crystal

For *every* calcium ion, there are ________fluoride ions.

We write the ***formula unit*** for calcium fluoride as:

$$CaF_2$$

Understanding Check:
Write the ***formula unit*** for the compound formed by combining magnesium and chloride ions.

$$Mg_{\square}Cl_{\square}$$

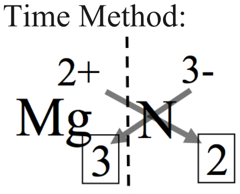
Understanding Check:
Write the ***formula unit*** for the compound formed by combining potassium and oxide ions.

$$K_{\square}O_{\square}$$

Understanding Check:
Write the ***formula unit*** for the compound formed by combining magnesium and nitride ions.

$$Mg_{\square}N_{\square}$$

Dr. Zoval's Caveman Style, Works Every Time Method:



The Criss-Cross Method

Formula Units

Write the formula for the ionic compound formed between each of the following pairs of ions:

Cu^{+} and O^{2-}

Fe^{3+} and S^{2-}

Cu^{2+} and Cl^{-}

Mg^{2+} and O^{2-}

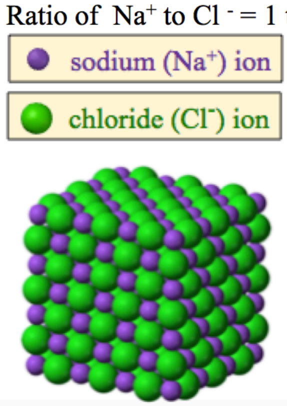
Sn^{4+} and S^{2-}

V^{3+} and Cl^{-}

Formula Unit vs. Molecular Formula

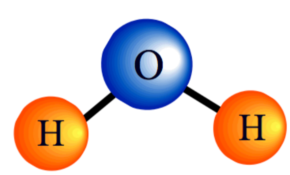
Formula Unit =
Lowest **RATIO** of ions

Example: NaCl
Ratio of Na^{+} to Cl^{-} = 1 to 1



Molecular Formula =
Actual **number** of atoms

Example: H_2O
two hydrogen atoms and
one oxygen atom



Naming Ionic Compounds

Educational Goals:

Given the **name** of an ***ionic compound***, be able to write the **formula unit**.
Given the **formula unit** of an ***ionic compound***, be able to write the **name**.

Method for Writing Formula Units for Ionic Compounds

1) Write the symbol of the first ion (the cation) in the compound's name, then the symbol of the second ion (the anion) in the compound's name.

2) Indicate the **ratio** of the ions in the compound using ***subscripts*** after each ion.

The ratio of the ions is deduced by *balancing the charges* of the ions so that the total charge in the crystal is equal to **zero**.

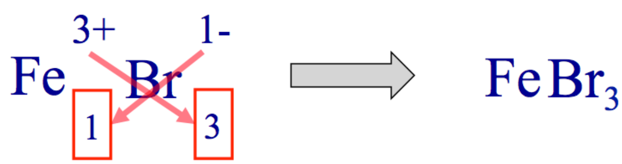
- We find the ion's charge from its position on the periodic table or, for polyatomic ions, we look it up in a table.
- You will know the charge for the metals t*hat occur with various charges* because the charge will be written in the compound's name in Roman numerals.

For ________________ ***ions***:
When the ***subscript*** for a **polyatomic ion** is ***greater than 1***, the polyatomic ion formula is written in parenthesis and the subscript is written after/outside of the parenthesis.

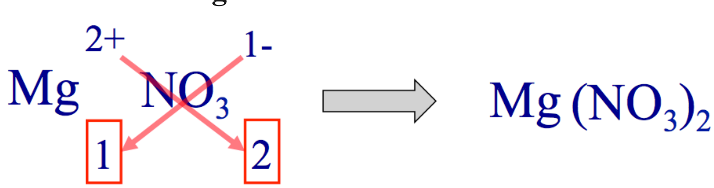
Example: Write the ***formula unit*** for **iron(III) bromide**.

- You will know the **charge** for the metals t*hat occur with various charges* because the charge will be written in the compound's name in **Roman numerals**.

iron(III) bromide



Example: Write the ***formula unit*** for **magnesium nitrate**.



- **For polyatomic ions**:
  When the ***subscript*** for a **polyatomic ion** is ***greater than 1***, the polyatomic ion formula is written in parenthesis and the subscript is written after/outside of the parenthesis.

Understanding Check: Write the *formula unit* for each of the following compounds:

a. sodium bicarbonate
b. sodium fluoride
c. iron(III) chloride
d. sodium carbonate
e. copper(II) sulfate
f. magnesium hydroxide

Method for Writing the Names of Ionic Compounds

1. Write the ____________ name first, then the _____________ name.
    - Monoatomic ***anions*** (anions composed of one element) use the "ide" suffix.
    - We get the names of *polyatomic ions* from the polyatomic ion table.
2. If the cation is one of the *metals with various charges*, write the charge using parenthesis and Roman numerals after the metal's name.

Example: Name the following compound: $\mathbf{MgCl_2}$

Name: **magnesium chloride**

Example: Name the following compound: $\mathbf{CuBr_2}$

- What *must* the charge of the copper ion be? **2+**

Name: **copper(II) bromide**

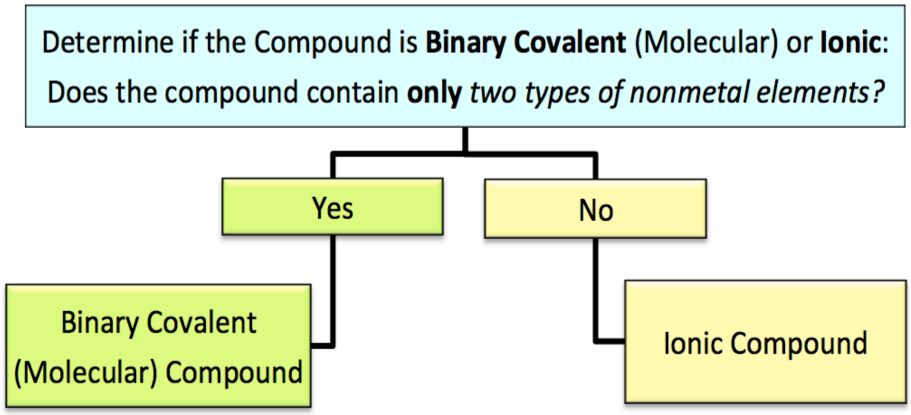
Complete the names of the following ionic compounds with variable charge metal ions:

$FeBr_2$	iron(__) bromide
$CuCl$	copper(__) chloride
SnO_2	______(__) _______
Fe_2O_3	___________________

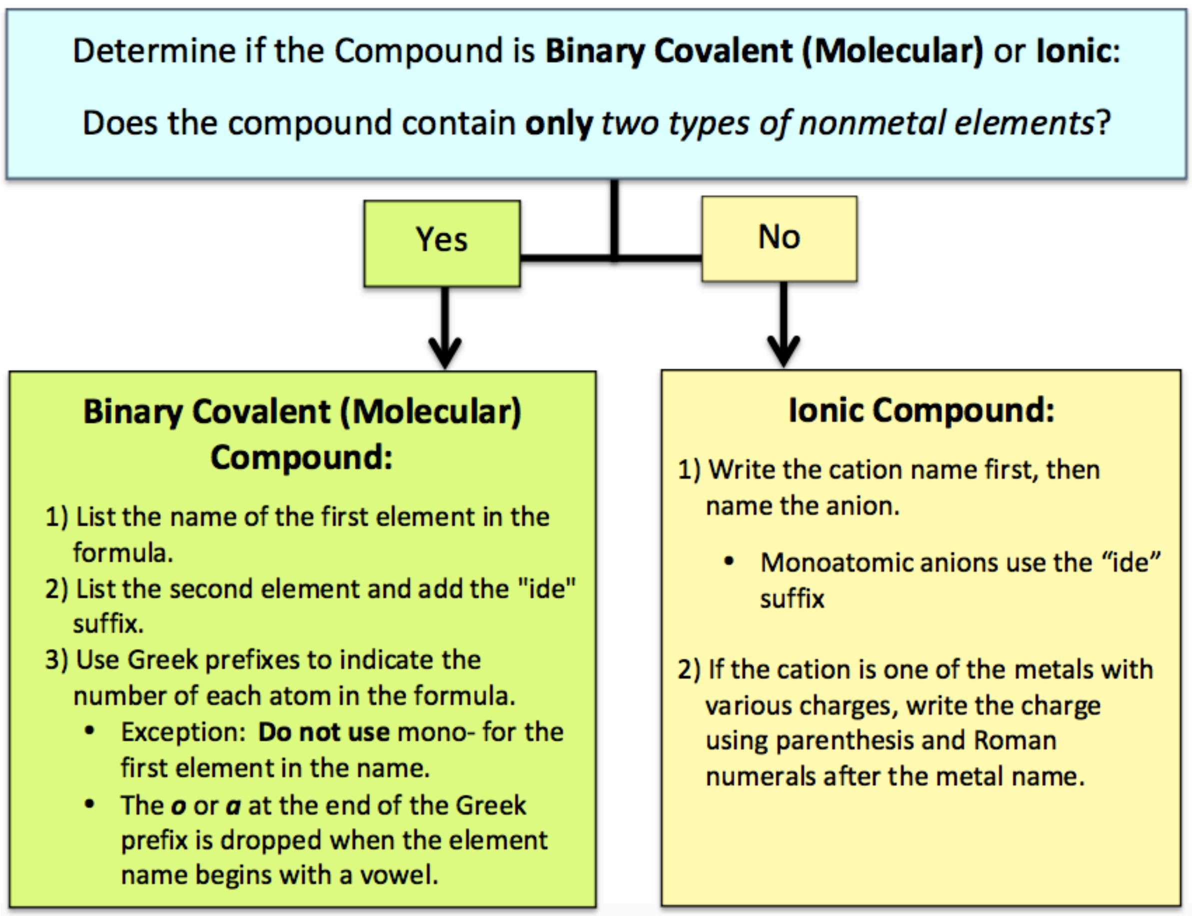
Name the following ionic compounds:

$NaCl$	_____________________
ZnI_2	_____________________
Al_2O_3	_____________________

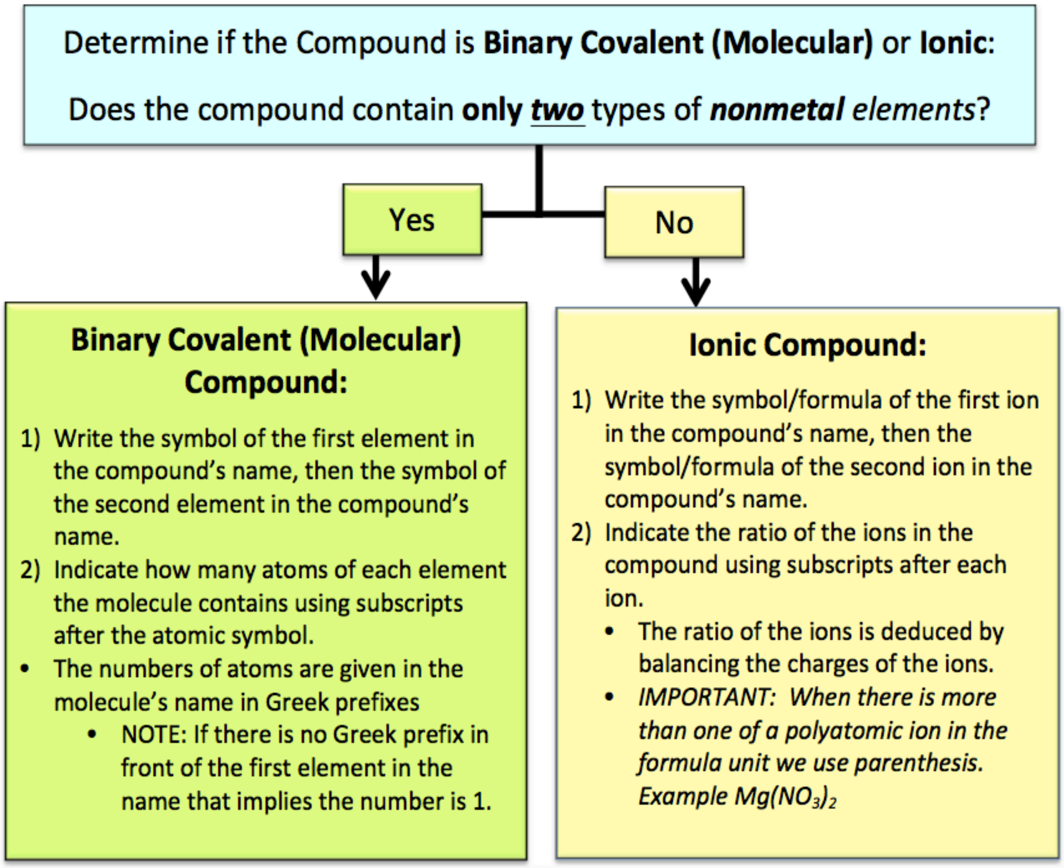
Naming Compound Summary



Given the Molecular Formula, Write the Name



Given the Name, Write the Molecular Formula



Molar Mass of Compounds

In this video, you will learn how to calculate the **molar mass** of a compound and how to use the molar mass of a compound to do ***mole-mass conversions***.

1) Molar Mass of **Covalent Compounds** (**Molecules**)
2) Molar Mass of **Ionic Compounds**

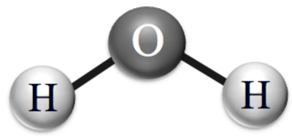
Molar Mass of Covalent Compounds (Molecules)

The **molar mass** of a ___________________ tells us the mass (grams) of **1 mole** of the ***molecules***.

- The *molar mass* of a molecule is also called the **molecular mass**.

To calculate the *molar mass* of a **molecule** we **add up** the ***atomic molar masses*** of **all** ____________ in the molecule.

Example: Let's calculate the molar mass of H_2O.



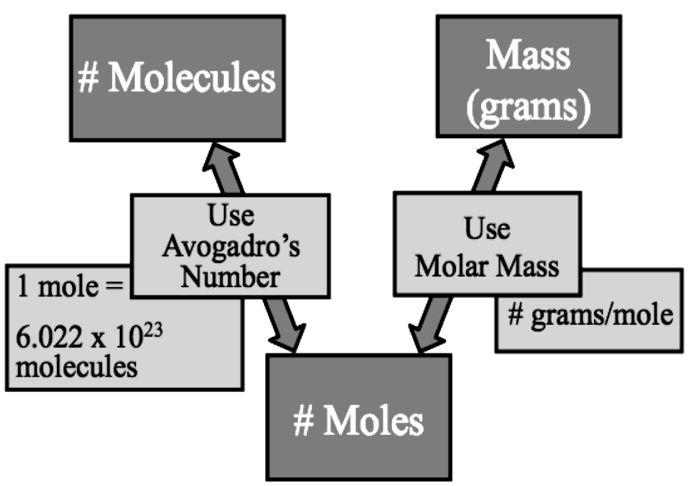
Atom	# of Atoms	Atomic Molar Mass	Total
oxygen	1	x 16.00 g/mole	16.00 g/mole
hydrogen	2	x 1.01 g/mole	2.02 g/mole
		Molar Mass of H_2O =	**18.02 g/mole**

One mole of H_2O (6.022 x 10^{23} molecules) } has a mass of 18.02 grams

Understanding Check: Calculate the molar mass of CH_4 (methane).

Mass-Mole-Molecules Conversions

Note that, as in the case of atoms, the molar mass of a compound is the ***relationship*** between *moles* and *mass (grams)*, therefore we can **convert** between moles and grams of compounds.



Example: How many *grams* of CH_4 is contained in 3.65 *moles*?

Use the molar mass to write an *equivalence statement*:

- **1 mole CH_4 = 16.05 grams**

The equivalence statements can be written as conversion factors:

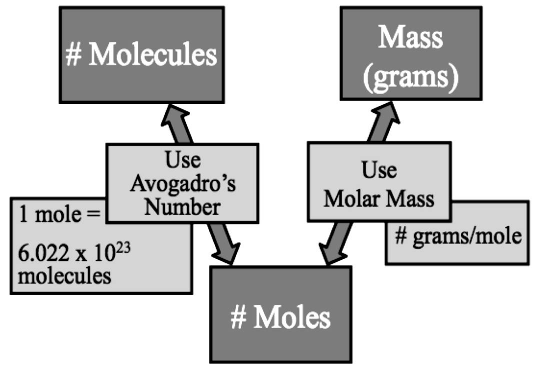
$$\left(\frac{1 \text{ mole } CH_4}{16.05 \text{ grams}}\right) \xleftrightarrow{\text{Conversion Factors}} \left(\frac{16.05 \text{ grams}}{1 \text{ mole } CH_4}\right)$$

$$3.65\ \cancel{\text{moles } CH_4} \times \frac{\textbf{16.05 grams } CH_4}{1\ \cancel{\text{mole } CH_4}} = 58.6 \text{ grams } CH_4$$

You have just learned how to convert between moles and mass of a compound and vice versa.

We do a **two-step calculation** to convert between **mass** and number of **molecules**.

We can *convert between molecules and moles* since **Avogadro's Number applies to molecules**; one mole of a molecular compound contains 6.022 x 10^{23} molecules.



You try one: How many H_2O ***molecules*** are contained in 237 grams?

Molar Mass of Ionic Compounds

When using the **molar mass** of *ionic compounds*, we calculate the mass of a compound based on the number of each ion as it appears in the formula unit.

- For this reason, the *molar mass of an ionic compound* is also called __________________ **mass**.

Example: The molar mass of sodium chloride (NaCl)

The **formula unit** for *sodium chloride* is **NaCl** because there is a 1:1 ratio of sodium ions to chloride ions in the crystal.

One **mole** of sodium chloride contains ***one*** **mole** of *sodium ions* and ***one*** **mole** of *chloride ions*.

Although **ions** have *extra* or *missing* elections, their molar masses are calculated by adding the *atomic molar masses* of the elements they contain.

- The reason we can do this is because the mass of electrons is negligible compared to the mass of protons and neutrons.

Ion	# of ions in the Formula Unit	Molar Mass of ion	Total
Sodium	1	x 22.99 g/mole	= 22.99 g/mole
Chloride	1	x 35.45 g/mole	= 35.45 g/mole
Molar Mass (Formula Mass) of NaCl			**= 58.44 g/mole**

Example: What is the molar mass of iron(II) phosphate, $Fe_3(PO_4)_2$?

One mole of iron(II) phosphate contains ***three*** moles of *iron(II) ions* and ***two*** moles of *phosphate ions*.

three moles of *iron(II) ions*

Fe_3

The molar mass of **each** *iron(II)* ion is: 55.85 g/mole.

two moles of *phosphate ions*

$(PO_4)_2$

each phosphate ion contains:

- one mole of phosphorus
- four moles of oxygen

The molar mass of **each** *phosphate ion* is: **94.97 g/mole**.

The molar mass (or formula mass) is calculated by adding the molar masses of the ions:

Ion	# of Ions in the Formula Unit	Molar Mass of ion	Total
Iron(II)	3	x 55.85 g/mole	= 167.55 g/mole
Phosphate	2	x 94.97 g/mole based on: one phosphorus **and** four oxygens ***per ion***	= 189.94 g/mole
Molar Mass (Formula Mass) of $Fe_3(PO_4)_2$			**= 357.49 g/mole**

An Alternative Method:

Fe_3

three **moles** of *iron(II) ions*

$(PO_4)_2$

two **moles** of *phosphate ions* contain:

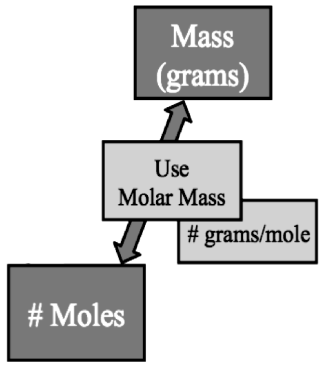
- ***two*** moles of phosphorous
- ***eight*** (**2** x 4) moles of oxygen

Three moles of **Fe**:	**3** x 55.85 g/mole =	**167.55** g/mole
Two moles of **P**:	**2** x 30.97 g/mole =	**061.94** g/mole
Eight moles of **O**:	**8** x 16.00 g/mole =	**128.00** g/mole
	The molar mass of $Fe_3(PO_4)_2$ is	**357.49** g/mole

Understanding Check: What is the molar mass of magnesium nitrate, $Mg(NO_3)_2$?

Mole-Mass Conversions for Ionic Compounds

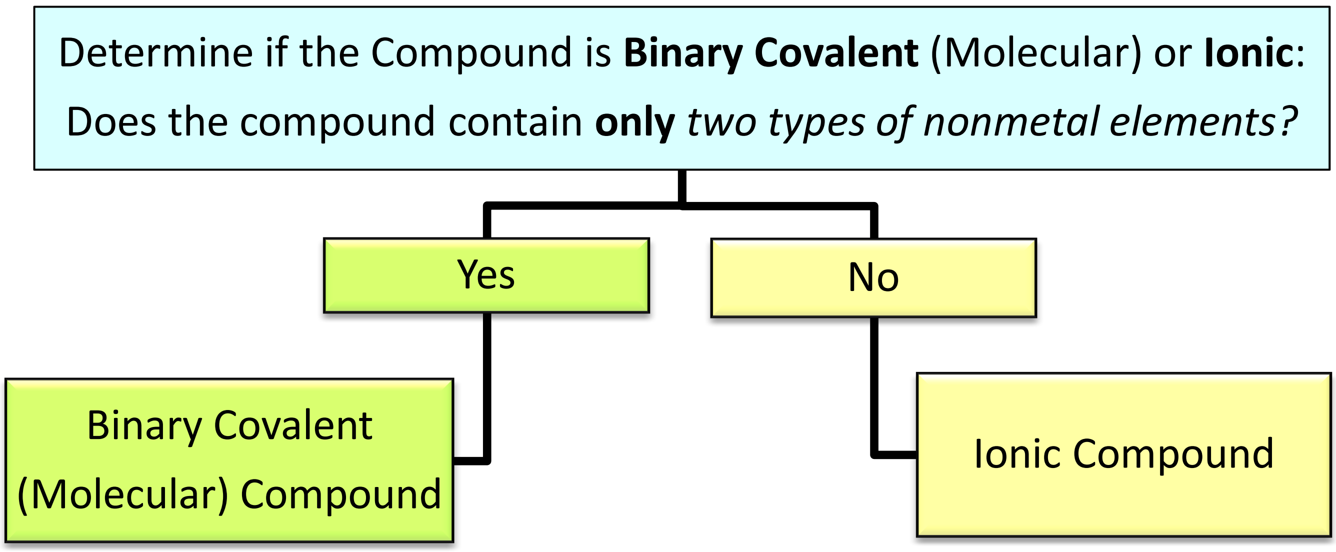
Mole-Mass conversions for ionic compounds are done *exactly* as we did for covalent compounds; use the molar mass as a conversion factor.



You Try One: What is the **mass** (grams) of 4.95 **moles** of $Mg(NO_3)_2$?

Naming Compounds Tutorial and Worksheet

Since we use different methods in naming binary covalent (molecular) compounds and ionic compounds, the **first step** in naming or writing the formula of a compound is to **determine which of the 2 compound classes it belongs**. This can be done as follows:



Binary covalent compounds will contain **only** *two types of non-metal elements*. There may be more than one of each element. For example CO_2 contains just two types of elements, carbon and oxygen. We will discuss naming covalent compounds that contain more than two types of elements, like glucose $C_6H_{12}O_6$, in later chapters.

Once it is determined that the compound is **ionic** or **covalent**, the student can be asked to do either of the following:

1) Given the **name** of the compound, write the **formula**.

<u>Or</u>

2) Given the **formula** of the compound, write the **name**.

In this tutorial we will review the process for achieving these 2 objectives and practice with some worksheet problems. First, we will review and practice how to write formulas for compounds when given the compound's name. Second, we will review and practice how to write the name of a compound when given the compound's formula.

Given the Name of the Compound, Writing Formulas for Compounds

Determine if the Compound is **Binary Covalent (Molecular)** or **Ionic**:

Does the compound contain **only _two_** types of ***nonmetal*** *elements*?

Yes

No

Binary Covalent (Molecular) Compound:

1) Write the symbol of the first element in the compound's name, then the symbol of the second element in the compound's name.
2) Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.

- The numbers of atoms are given in the molecule's name in Greek prefixes
    - NOTE: If there is no Greek prefix in front of the first element in the name that implies the number is 1.

Ionic Compound:

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.
2) Indicate the ratio of the ions in the compound using subscripts after each ion.
    - The ratio of the ions is deduced by balancing the charges of the ions.
    - *IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. Example $Mg(NO_3)_2$*

Writing the Formulas of Ionic Compounds

Example: Write the formula for **calcium bromide**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.

   Ca Br

2) Indicate the ratio of the ions in the compound using subscripts after each ion.
   - This step involves filling in the subscripts boxes as we did in the lecture:

     $Ca_{\boxed{\ }}Br_{\boxed{\ }}$

   - The ratio of the ions is deduced by **balancing the charges** of the ions.
     - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
     - We find the ion's charge from its position on the periodic table <u>**or**</u> we look it up in a table in the case of polyatomic ions.
     - Transition metal with varying charges will be written in the compound name in Roman numerals.

   - First, temporarily write the charge of each ion above the ion's symbol.

     $$\overset{2+}{Ca}_{\boxed{\ }}\overset{1-}{Br}_{\boxed{\ }}$$

   - Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **two** bromide ions, each has a charge of (1-) to cancel the (2+) charge of the calcium ion:
     - 2(-1) + (+2) = 0 zero total charge.

     $$\overset{2+}{Ca}_{\boxed{1}}\overset{1-}{Br}_{\boxed{2}} \quad = \quad CaBr_2$$

     - We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

     $$\overset{2+}{Ca}_{\boxed{1}}\overset{1-}{Br}_{\boxed{2}} \quad = \quad CaBr_2$$

     - Note, we do not leave the charges written above the symbols in the completed formula.

IMPORTANT: When there is more than one of a polyatomic ion in the formula, we use parenthesis.

- Not applicable in this example since there are no polyatomic ions in calcium bromide.

Examples: Writing the Formulas of Ionic Compounds

Write the formula for **magnesium nitrate**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.
    - When you see a polyatomic ion (nitrate), look up the formula and charge in the table of polyatomic ions.

$$Mg \;\; NO_3$$

2) Indicate the ratio of the ions in the compound using subscripts after each ion.
    a. This step involves filling in the subscripts boxes as we did in the lecture:

$$Mg_{\boxed{\ }}NO_{3\boxed{\ }}$$

- The ratio of the ions is deduced by **balancing the charges** of the ions.
    - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
    - We find the ion's charge from its position on the periodic table **<u>or</u>** we look it up in a table in the case of polyatomic ions.
    - Transition metal with varying charges will be written in the compound name in Roman numerals.

- First, temporarily write the charge of each ion above the ion's symbol.

$$\overset{2+}{Mg}_{\boxed{\ }}\overset{1-}{NO_3}_{\boxed{\ }}$$

- Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **two** nitrate ions, each has a charge of (1-) to cancel the (2+) charge of the magnesium ion:
    - 2(-1) + (+2) = 0 zero total charge.

$$\overset{2+}{Mg}_{\boxed{1}}\overset{1-}{NO_3}_{\boxed{2}}$$

- We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

$$\overset{2+}{Mg}_{\boxed{1}}\overset{1-}{NO_3}_{\boxed{2}}$$

*IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. There are **2 ions** of nitrate in magnesium nitrate*

$$Mg_{\boxed{1}}NO_{3\boxed{2}} = Mg(NO_3)_2$$

In compound where there is just **one formula unit** of a polyatomic ion, no parenthesis are needed. An example of this is **sodium nitrate: $NaNO_3$**

Examples: Writing the Formulas of Ionic Compounds

Write the formula for **iron(II) phosphate**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.
    - When you see a polyatomic ion (phosphate in this case), look up the formula and charge in the table of polyatomic ions.

$$Fe\ \ PO_4$$

2) Indicate the ratio of the ions in the compound using subscripts after each ion.
    b. This step involves filling in the subscripts boxes as we did in the lecture:

$$Fe_{\square}\,PO_{4\square}$$

- The ratio of the ions is deduced by **balancing the charges** of the ions.
    - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
    - We find the ion's charge from its position on the periodic table **<u>or</u>** we look it up in a table in the case of polyatomic ions.
    - **Transition metal with varying charges will be written in the compound name in Roman numerals.**
        - In this example, now we know the charge on **the Fe ion is 2+**

- First, temporarily write the charge of each ion above the ion's symbol.

$$\overset{2+}{Fe_{\square}}\,\overset{3-}{PO_{4\square}}$$

- Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **two** phosphate ions, each has a charge of (3-) and three Fe^{2+} ions to balance the charge:
    - 2(-3) + 3(-2) = 0 zero total charge.

$$\overset{2+}{Fe_{\boxed{3}}}\,\overset{3-}{PO_{4\boxed{2}}}$$

- We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

$$\overset{2+}{Fe_{\boxed{3}}}\,\overset{3-}{PO_{4\boxed{2}}}$$

*IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. There are **2 ions** of phosphate in iron(II)phosphate.*

$$Fe_{\boxed{3}}\,PO_{4\boxed{2}} = Fe_3(PO_4)_2$$

Examples: Writing the Formulas of Ionic Compounds

Write the formula for **barium sulfide**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.

   Ba S

2) Indicate the ratio of the ions in the compound using subscripts after each ion.
   - This step involves filling in the subscripts boxes as we did in the lecture:

     $Ba_{\boxed{}}S_{\boxed{}}$
   - The ratio of the ions is deduced by **balancing the charges** of the ions.
     - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
     - We find the ion's charge from its position on the periodic table **<u>or</u>** we look it up in a table in the case of polyatomic ions.
     - Transition metal with varying charges will be written in the compound name in Roman numerals.

   - First, temporarily write the charge of each ion above the ion's symbol.

     $\overset{2+}{Ba}_{\boxed{}}\overset{2-}{S}_{\boxed{}}$

   - Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **one** sulfide ion, with a charge of (2-) to cancel the (2+) charge of the barium ion:
     - (-2) + (+2) = 0 zero total charge.

       $$\overset{2+}{Ba}_{\boxed{1}}\overset{2-}{S}_{\boxed{1}} = BaS$$

     - We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

       $$\overset{2+}{Ba}_{\boxed{2}}\overset{2-}{S}_{\boxed{2}} = BaS$$

     - Note, the subscripts in ionic compound represent the ratio in which large numbers of anions and cations combine to form the ionic compounds. Since we want the **lowest ratio**: we use 1:1, since 2:2 = 1:1

       $$\cancel{Ba_2S_2} = BaS$$

Write the formula for the following ionic compounds: (see next page for key)

sodium bicarbonate ________________

sodium fluoride ________________

iron (III) chloride ________________

sodium carbonate ________________

copper (II) sulfate ________________

magnesium hydroxide ________________

barium nitrate ________________

lithium sulfate ________________

magnesium chloride ________________

silver nitrate ________________

aluminum sulfate ________________

calcium hydroxide ________________

calcium sulfate ________________

mercury (II) nitrate ________________

lead (IV) nitrate ________________

magnesium iodide ________________

sodium nitride ________________

Practice Problems KEY

sodium bicarbonate $NaHCO_3$

sodium fluoride   NaF

iron (III) chloride $FeCl_3$

sodium carbonate Na_2CO_3

copper (II) sulfate $CuSO_4$

magnesium hydroxide $Mg(OH)_2$

barium nitrate $Ba(NO_3)_2$

lithium sulfate Li_2SO_4

magnesium chloride $MgCl_2$

silver nitrate $AgNO_3$

aluminum sulfate $Al_2(SO_4)_3$

calcium hydroxide $Ca(OH)_2$

calcium sulfate $CaSO_4$

mercury (II) nitrate $Hg(NO_3)_2$

lead (IV) nitrate $Pb(NO_3)_4$

magnesium iodide MgI_2

sodium nitride Na_3N

Writing the Formulas of Covalent Compounds

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second element in the compound's name.
2) Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.
    - The numbers of atoms are given in the molecule's name in Greek prefixes
    - NOTE: If there is no Greek prefix in front of the first element in the name, that means the number is 1.

Example: Write the formula of **dinitrogen tetrafluoride**.

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second element in the compound's name.

$$N \;\; F$$

2) Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.

$$N_{\square}F_{\square}$$

- The numbers of atoms are given in the molecule's name in Greek prefixes.
    - **di**nitrogen **tetra**fluoride
    - see your chapter 3 notes for a list of the Greek prefixes

$$N_2F_4$$

- **NOTE**: If there is no Greek prefix in front of the first element in the name, then the number is 1.
    - Example carbon tetrachloride = CCl_4

Example: Write the formula of **carbon disulfide**.

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second element in the compound's name.

$$C \;\; S$$

2) Indicate how many atoms of each element the molecule contains using subscripts after the atomic symbol.

$$C_{\square}S_{\square}$$

- The numbers of atoms are given in the molecule's name in Greek prefixes.
    - carbon **di**sulfide
    - see your chapter 3 notes for a list of the Greek prefixes

$$C_1S_2 \quad = CS_2$$

- **NOTE**: If there is no Greek prefix in front of the first element in the name, then the number is 1.

Write the formulas for the following covalent compounds:

See next page for KEY

a. disulfur tetrafluoride __________________________

b. carbon trioxide __________________________

c. nitrogen pentoxide __________________________

d. nitrogen tribromide __________________________

e. dinitrogen heptachloride __________________________

f. carbon tetrachloride__________________________

g. hydrogen monochloride__________________________

h. trihydrogen monophosphide __________________________

i. dihydrogen monoxide __________________________

KEY

a. disulfur tetrafluoride S_2F_4

b. carbon trioxide CO_3

c. nitrogen pentoxide NO_5

d. nitrogen tribromide NBr_3

e. dinitrogen heptachloride N_2Cl_7

f. carbon tetrachloride CCl_4

g. hydrogen monochloride HCl

h. trihydrogen monophosphide H_3P

i. dihydrogen monoxide H_2O

Given the Formulas for Compounds, Write the Name

Determine if the Compound is **Binary Covalent (Molecular)** or **Ionic**:

Does the compound contain **only** *two types of nonmetal elements*?

Yes

No

Binary Covalent (Molecular) Compound:

1) List the name of the first element in the formula.
2) List the second element and add the "ide" suffix.
3) Use Greek prefixes to indicate the number of each atom in the formula.
   - Exception: **Do not use** mono- for the first element in the name.
   - The ***o*** or ***a*** at the end of the Greek prefix is dropped when the element name begins with a vowel.

Ionic Compound:

1) Write the cation name first, then name the anion.
   - Monoatomic anions use the “ide” suffix

2) If the cation is one of the metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.

Writing the Names of Ionic Compounds

Example: Write the name for $\mathbf{CaBr_2}$

1) Write the cation name first, then name the anion.
   - monoatomic anions use the "ide" suffix

   calcium bromide

2) If the cation is one of the transition metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.
   - Not necessary here, there is not a transition metal present

Example: Write the name for $\mathbf{Mg(NO_3)_2}$

1) Write the cation name first, then name the anion.
   - monoatomic anions use the "ide" suffix
     - Here we notice that the anion is a **polyatomic ion**. Get the name from the polyatomic ion table (in your notes or textbook). *You will be given a copy of the polyatomic ion table on your exams*.
     - **Do not** change the suffix to "ide" with polyatomic ions:

   magnesium nitrate

2) If the cation is one of the transition metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.
   - Not necessary here, there is not a transition metal present

Writing the Names of Ionic Compounds

Example: Write the name for **CuF_2**

1) Write the cation name first, then name the anion.
   - monoatomic anions use the "ide" suffix

   copper fluoride

2) If the cation is one of the ***transition metals*** with various charges, write the **charge using parenthesis and Roman numerals** after the metal name.

   copper(?) fluoride

- We must figure out what the charge is on the copper, we can deduce the charge on the transition metal cations from the charge on the anions
  - Recall that the total charge for any compound must equal zero.
  - Since there are two fluorides, each with a charge of (1-) and there is only one copper, we can conclude that the charge on the copper must be (2+).
    - You can think of this as the reverse criss-cross! See chapter 3 notes for more details.

copper(II) fluoride

write the **charge in parenthesis in Roman numerals after the cation name**

Write the names of the following compounds:

See next page for key

$NaCl$ ____

$Fe_2(CO_3)_3$ ____

$Cu(OH)_2$ ____

$(NH_4)_2SO_4$ ____

$LiNO_3$ ____

$BaSO_4$ ____

$Mg(NO_3)_2$ ____

$AgCl$ ____

$Al(OH)_3$ ____

$CaSO_4$ ____

FeS ____

$FeCl_3$ ____

NaI ____

$MgCO_3$ ____

KEY

NaCl   sodium chloride

$Fe_2(CO_3)_3$ iron(III) carbonate

$Cu(OH)_2$ copper(II) hydroxide

$(NH_4)_2SO_4$ ammonium sulfate

$LiNO_3$ lithium nitrate

$BaSO_4$ barium sulfate

$Mg(NO_3)_2$ magnesium nitrate

AgCl   silver chloride

- (note: silver is one of the transition metals that only occurs as a (1+) ion)

$Al(OH)_3$ aluminum hydroxide

$CaSO_4$ calcium sulfate

FeS Iron(II) sulfide

$FeCl_3$ iron(III) chloride

NaI sodium iodide

$MgCO_3$ magnesium carbonate

Writing the Names of Covalent Compounds

1) List the name of the first element in the formula.

2) List the second element and add the –ide suffix.

3) Use Greek prefixes to indicate the number of each atom in the formula.
   - Exception: do not use mono- for the first element in the name.
   - The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel

Example: Write the name for $\mathbf{N_2S_4}$

1) List the name of the first element in the formula.

   nitrogen

2) List the second element and add the –ide suffix.

   nitrogen sulfide

3) Use Greek prefixes to indicate the number of each atom in the formula.
   - See your textbook or lecture notes for a table of the Greek prefixes.

   ____nitrogen _____sulfide

   dinitrogen tetrasulfide

   - Exception: do not use mono- for the first element in the name.
     - Not applicable in this example
   - The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
     - Not applicable in this example

Example: Write the name for $\mathbf{SO_3}$

1) List the name of the first element in the formula.

   sulfur

2) List the second element and add the –ide suffix.

   sulfur oxide

3) Use Greek prefixes to indicate the number of each atom in the formula.

   ____ sulfur _____ oxide

   sulfur trioxide

   - Exception: do not use **mono-** for the ***first*** element in the name.
     - NOTE, we did not write **mono**sulfur because of this rule!
   - The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
     - Not applicable in this example

Example: Write the name for $\mathbf{SO_2}$

1) List the name of the first element in the formula.

sulfur

2) List the second element and add the –ide suffix.

sulfur oxide

3) Use Greek prefixes to indicate the number of each atom in the formula.

____ sulfur _____ oxide

sulfur dioxide

- Exception: do not use **mono-** for the ***first*** element in the name.
  - NOTE, we did not write **mono**sulfur because of this rule!
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
  - Not applicable in this example

Example: Write the name for **CO**

1) List the name of the first element in the formula.

carbon

2) List the second element and add the –ide suffix.

carbon oxide

3) Use Greek prefixes to indicate the number of each atom in the formula.

____ carbon _____ oxide

carbon monoxide

- Exception: do not use **mono-** for the ***first*** element in the name.
  - NOTE, we did not write **monocarbon** because of this rule!
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
  - NOTE, we did not write **monooxide** because of this rule!

Write the names of the following compounds:

See next page for key

a. Br_2I_4 ____

b. P_5F_8 ____

c. NO_5 ____

- Remember: The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel

d. NBr_3 ____

e. N_2O_5 ____

f. $BrCl_3$ ____

g. H_2S ____

h. N_2O ____

KEY

a. Br_2I_4 dibromine tetriodide

b. P_5F_8 pentaphosphorus octafluoride

c. NO_5 nitrogen pentoxide
   - The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
     - NOTE, we did not write **pent<u>ao</u>xygen** because of this rule!

d. NBr_3 nitrogen tribromide

e. N_2O_5 dinitrogen pentoxide

f. $BrCl_3$ bromine trichloride

g. H_2S dihydrogen monosulfide

h. N_2O dinitrogen monoxide

Molar Mass Worksheet and Key

For the following compounds, write the chemical formula and determine the molar mass. Write the units!

water

sodium carbonate

carbon dioxide

barium nitrate

sodium chloride

hydrogen monochloride

calcium hydroxide

sulfuric acid (H_2SO_4)

potassium permanganate ($KMnO_4$)

acetic acid ($C_2H_4O_2$)

SEE NEXT PAGE FOR KEY

Molar Mass Worksheet Key

Water (H_2O)

Element	Number	MM
H	2	1.01 g/mole
O	1	16.00 g/mole

Total MM = 18.02 g/mole

carbon dioxide, CO_2

Element	Number	MM	
O	2	16.00	g/mole
C	1	12.01	g/mole

Total MM = 44.01 g/mole

sodium chloride, NaCl

Element	Number	MM	
Cl	1	35.45	g/mole
Na	1	22.99	g/mole

Total MM = 58.44 g/mole

calcium hydroxide, $Ca(OH)_2$

Element	Number	MM	
H	2	1.01	g/mole
O	2	16.00	g/mole
Ca	1	40.08	g/mole

Total MM = 74.10 g/mole

potassium permanganate ($KMnO_4$)

Element	Number	MM	
O	4	16.00	g/mole
Mn	1	54.94	g/mole
K	1	39.10	g/mole

Total MM = 158.04 g/mole

sodium carbonate, Na_2CO_3

Element	Number	MM	
O	3	16.00	g/mole
Na	2	22.99	g/mole
C	1	12.01	g/mole

Total MW = 105.99 g/mole

barium nitrate, $Ba(NO_3)_2$

Element	Number	MM	
O	6	16.00	g/mole
Ba	1	137.33	g/mole
N	2	14.01	g/mole

Total MM = 261.35 g/mole

hydrogen monochloride, HCl

Element	Number	MM	
Cl	1	35.45	g/mole
H	1	1.01	g/mole

Total MM = 36.46 g/mole

sulfuric acid (H_2SO_4)

Element	Number	MM	
S	1	32.07	g/mole
H	2	1.01	g/mole
O	4	16.00	g/mole

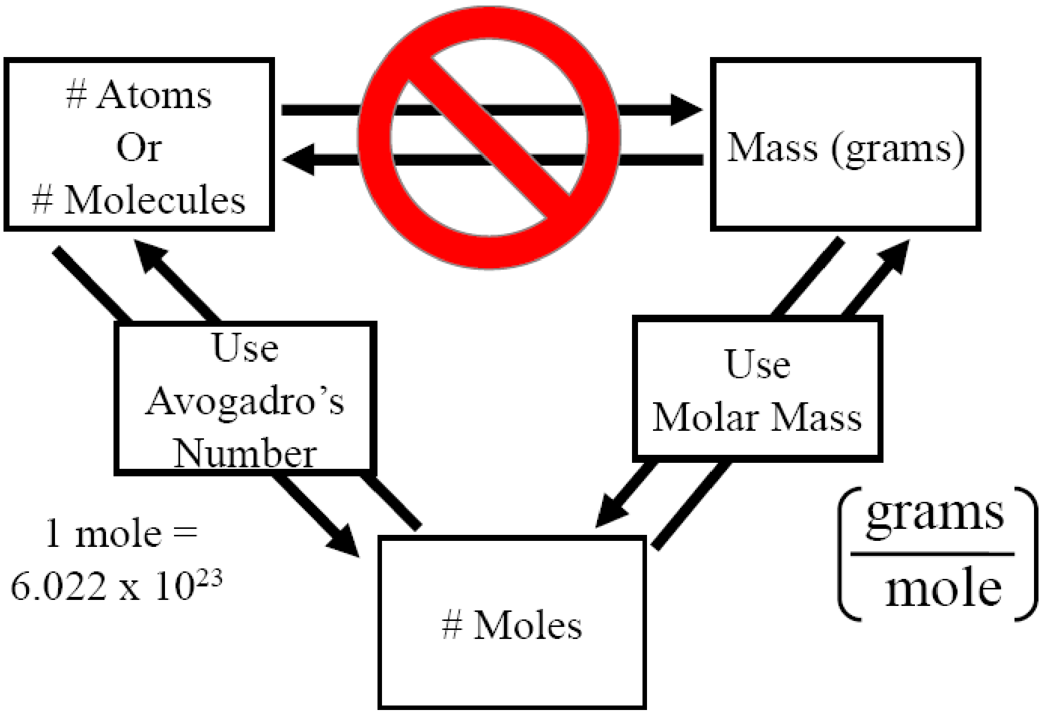
Total MM = 98.09 g/mole

acetic acid ($C_2H_4O_2$)

Element	Number	MM	
H	4	1.01	g/mole
O	2	16.00	g/mole
C	2	12.01	g/mole

Total MM = 60.06 g/mole

Moles, Molecules, and Grams Worksheet and Key



1) How many moles are there in 24.0 grams of FeF_3?

2) How many moles are there in 458 grams of Na_2SO_4?

3) How many grams are there in 2.30×10^{24} atoms of silver?

4) How many grams are there in 7.40 moles of $AgNO_3$?

SEE LAST PAGE FOR KEY

5) How many grams are there in 7.50×10^{23} molecules of H_2SO_4?

6) How many molecules are there in 122 grams of NO_2?

7) How many grams are there in 9.40×10^{25} molecules of H_2?

8) How many molecules are there in 237 grams of CCl_4?

9) How many molecules are there in 2.30 grams of NH_3?

10) How many grams are there in 3.30×10^{23} molecules of N_2I_6?

11) How many moles are there in 2.00×10^{19} molecules of CCl_4?

12) How many grams are there in 1.00 x 10^{24} molecules of BCl_3?

13) How many grams are there in 4.50 moles of $Ba(NO_2)_2$?

14) How many molecules are there in 9.34 grams of water?

Moles, Molecules, and Grams Worksheet – Answer Key

1) How many moles are there in 24.0 grams of FeF_3? **.213 moles**
2) How many moles are there in 458 grams of Na_2SO_4? **3.22 moles**
3) How many grams are there in 2.30 x 10^{24} atoms of silver? **412 grams**
4) How many grams are there in 7.40 moles of $AgNO_3$? **1260 grams (note:3 significant figures)**
5) How many grams are there in 7.50 x 10^{23} molecules of H_2SO_4? **122 grams**
6) How many molecules are there in 122 grams of NO_2? **1.60 x 10^{24} molecules**
7) How many grams are there in 9.40 x 10^{25} molecules of H_2? **315 grams**
8) How many molecules are there in 237 grams of CCl_4? **9.28 x 10^{23} molecules**
9) How many molecules are there in 2.30 grams of NH_3? **8.13 x 10^{22} molecules**
10) How many grams are there in 3.30 x 10^{23} molecules of N_2I_6? **433 grams**
11) How many moles are there in 2.00 x 10^{19} molecules of CCl_4? **3.32 x 10^{-5}moles**
12) How many grams are there in 1.00 x 10^{24} molecules of BCl_3? **195 grams**
13) How many grams are there in 4.50 moles of $Ba(NO_2)_2$? **1030 grams (3 significant figures)**
14) How many molecules are there in 9.34 grams of water? **3.12 x 10^{23} molecules**

Chapter 4 Lecture Notes

Educational Goals

1. Given the formula of a diatomic or small molecule, draw the **line bond structure.**
2. Understand and construct **condensed** and **skeletal structural formulas** given the line bond structures and vice versa.
3. Given the structural formula, determine the **formal charge** of **O** *and* **N** *atoms*.
4. Given the line bond structure of a ***small molecule***, predict the **molecular shape** and **bond angle(s)**.
5. Given the structure of a ***large molecule***, predict the **bond angle(s)** around any atom.
6. Define **electronegativity** and explain its relationship to *polar covalent bonds*.
7. Give a simple rule that can be used to predict whether or not a ***covalent bond*** is **polar**.
8. Classify diatomic, small, and large *molecules* as **polar** or **nonpolar**.
9. Describe, compare, and contrast the five **noncovalent interactions**.

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Review: The Octet Rule in the Formation of Molecules

Octet Rule: *Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an ____________ of electrons in its outermost shell (n).*

Exception to the octet rule ________________ and ______________.

- *Hydrogen* and *helium* have filled outer shells with just __________ electrons.

Covalent bonding results from the ***sharing*** of electron pairs between two atoms.

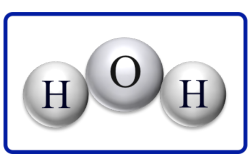
Method for Drawing Line Bond Structures

Step 1: Count the total number of __________________ __________________ from all the atoms in the molecule.

Example: H_2O

2 H atoms $2 \times 1e^- = 2e^-$

1 O atom $1 \times 6e^- = \underline{6e^-}$



Total number of valence e- = 8e-

The line bond structure of H_2O will have 8 electrons

Step 2: Draw the "Skeleton Structure"

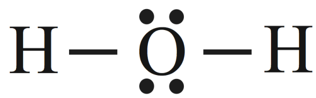
- Attach the atoms together with ____________ ____________in the most symmetric way possible.

$$H - O - H$$

Step 3: Subtract the number of electrons used to make the skeleton structure from the total number of valence electrons.

Total number of valence e^-	$8e^-$
Minus electrons used in skeleton	$-4e^-$
Electrons remaining **to be added** =	$4e$-

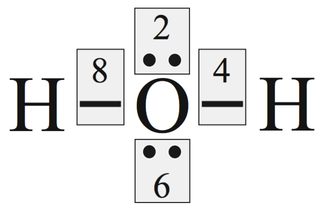
Step 4: Add the remaining electrons as ______________ ______________ as evenly as possible on all atoms except hydrogen.



Step 5: Check for ______________.

– Are there 8 electrons around all atoms (except hydrogen)?

If YES, you are finished!

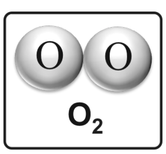


Step 6 (if needed): Use lone pairs to make ________________ or _________________ bonds until the octet rule is satisfied for *all atoms* in the molecule.

- Let's do a couple of examples to see how that works!
- We will do O_2, and N_2.

Example O_2

Step 1: Count the total number of valence electrons from all the atoms in the molecule.
- How many **valence electrons** in O_2?



2 O atoms $2 \times 6e^- = 12e^-$

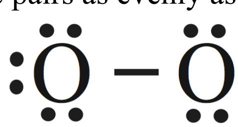
Total number of valence e^- $= 12e^-$

Step 2: Draw the "Skeleton Structure"
-Attach the atoms together with single bonds in the most symmetric way possible.

Step 3: Subtract the number of electrons used to make the skeleton structure from the total number of valence electrons.

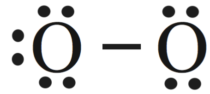
Total number of valence e^-	$12e^-$
Minus electrons used in skeleton	$-2e^-$
Electrons remaining *to be added* =	$10e$-

Step 4: Add the remaining electrons as lone pairs as evenly as possible on all atoms except hydrogen.



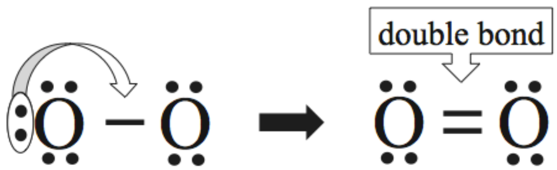
Step 5: Check for Octets

– Are there 8 electrons around all atoms (except hydrogen)?



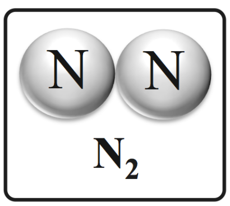
– If NO, use lone pairs to make double or triple bonds.

Step 6 (if needed): Use lone pairs to make **double** or **triple bond**s until the octet rule is satisfied for *all atoms* in the molecule.



Example N_2

Step 1: Count the total number of valence electrons from all the atoms in the molecule.



- How many **valence electrons** in N_2?

2 N atoms $2 \times 5e^- = 10e^-$

Total number of valence e^- $= 10e^-$

Step 2: Draw the "Skeleton Structure"

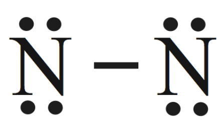
-Attach the atoms together with single bonds in the most symmetric way possible.

N — N

Step 3: Subtract the number of electrons used to make the skeleton structure from the total number of valence electrons.

Total number of valence e^-	$10e^-$
Minus electrons used in skeleton	$- 2e^-$
Electrons remaining *to be added* =	$8e$-

Step 4: Add the remaining electrons as lone pairs as evenly as possible on all atoms except hydrogen.

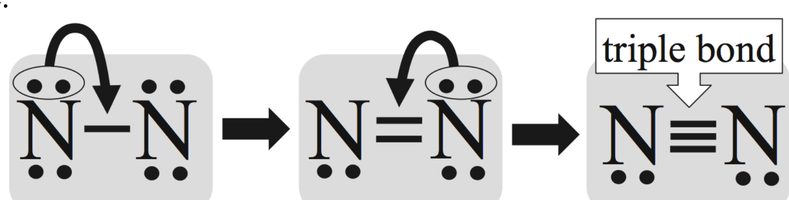


Step 5: Check for Octets

– Are there 8 electrons around all atoms (except hydrogen)?

– If NO, use lone pairs to make double or triple bonds.

Step 6 (if needed): Use lone pairs to make **double** or **triple bond**s until the octet rule is satisfied for *all atoms* in the molecule.



Draw the line bond structure for ammonia: **NH_3**

Step 1: Count the total number of valence electrons from all the atoms in the molecule.

Molecular formula of molecule ________

Atom	Number of atoms		Number of valence electrons per atom	Totals
N	1	x		=
H	3	x		=

Total # of electrons in line bond structure =_______

Step 2: Draw the "Skeleton Structure"
-Attach the atoms together with single bonds in the most symmetric way possible.

Central atom is:________________

Draw skeleton:

# of electrons used in skeleton = _________
(multiply # of bonds in skeleton by 2)

Step 3: Subtract the number of electrons used to make the skeleton structure from the total number of valence electrons.

Total # of electrons in structure _______
(from step 1 above)

# of electrons used in skeleton - _______
(from step 2 above)

Remaining # electrons to be added= _____

Step 4: First: Re-draw skeleton here:

Next, Add the remaining electrons as <u>lone pairs</u> as evenly as possible on all atoms except hydrogen.

Step 5: Check for Octets
Check the structure in step 4 for octets of electrons around each atom (except for hydrogen).

If the octet rule is satisfied, **you are done**.
If the octet rule is **<u>not</u>** satisfied, go the **Step 6**.

Step 6 (if needed): Use lone pairs to make **double** or **triple** bonds until the octet rule is satisfied for all atoms in the molecule.

Understanding Check: Draw the line bond structure for sulfur trioxide: **SO_3**

Step 1: Count the total number of valence electrons from all the atoms in the molecule.

Molecular formula of molecule ________

Atom	Number of atoms		Number of valence electrons per atom	Totals
				=
				=

Total # of electrons in line bond structure =_______

Step 2: Draw the "Skeleton Structure"
-Attach the atoms together with single bonds in the most symmetric way possible.

Central atom is:________________

Draw skeleton:

# of electrons used in skeleton = _________
(multiply # of bonds in skeleton by 2)

Step 3: Subtract the number of electrons used to make the skeleton structure from the total number of valence electrons.

Total # of electrons in structure _______
(from step 1 above)

\# of electrons used in skeleton - _______
(from step 2 above)

Remaining # electrons to be added= _____

Step 4: First: Re-draw skeleton here:

Next, Add the remaining electrons as <u>lone pairs</u> as evenly as possible on all atoms except hydrogen.

Step 5: Check for Octets
Check the structure in step 4 for octets of electrons around each atom (except for hydrogen).

If the octet rule is satisfied, **you are done**.
If the octet rule is **<u>not</u>** satisfied, go the **Step 6**.

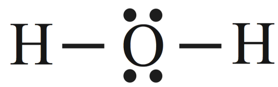
Step 6 (if needed): Use lone pairs to make **double** or **triple** bonds until the octet rule is satisfied for all atoms in the molecule.

Structural Formulas

A **molecular formula** shows us the number and types of atoms contained in a molecule.

- **Example** of a ***molecular formula***: H_2O

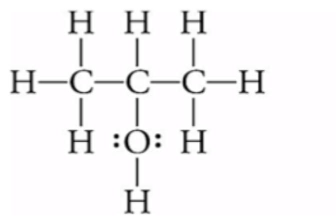
Drawings, such as *line bond structures*, that show the _________________of atoms within molecules are called _________________**formulas.**

- **Example** of a ***structural formula***: 

Structural Formulas vs. Molecular Formulas

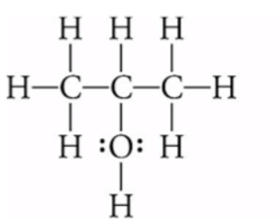
The **molecular formula** of *isopropyl alcohol* (often called rubbing alcohol) is C_3H_8O.

The **structural formula** of *isopropyl alcohol* is:

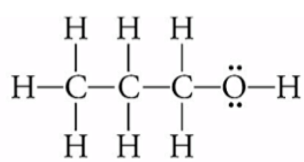


Which of these two types of formulas, the ***molecular formula*** or the ***structural formula***, do you think is more informative?

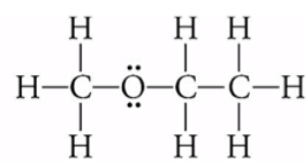
There are actually **three *completely different molecules*** that have this same molecular formula of $\mathbf{C_3H_8O}$:



isopropyl alcohol



1-propanol



ethyl methyl ether

Line bond structural formulas **explicitly** show all the information about how the atoms are connected and the presence of all single bonds, double bonds, triple bonds and lone pairs.

In practice, chemists often use one of two **short-cut** forms of structural formulas:

- _______________ ***structural formulas***

or

- _______________ ***structural formulas***

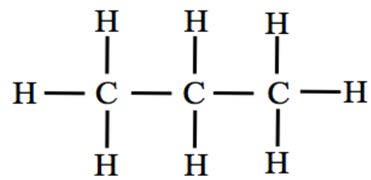
Condensed Structural Formulas (Condensed Structures)

When drawing **condensed structures**:

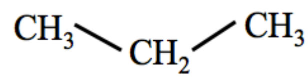
- Single bonds between carbon and hydrogen ***are omitted***.
  - **Example:** a carbon bonded to three hydrogen atoms is drawn as "$\mathbf{CH_3}$"
- Single bonds between carbons (C-C) ***can be*** omitted.
- Single bonds between oxygen or nitrogen ***and*** hydrogen ***can be*** omitted.
  - **Example:** an oxygen bonded to one hydrogen atom can be drawn as "**OH**."
- Bonds between ***all other*** pairs of atoms **must** be drawn.
- Lone pairs ***can be*** omitted.
- Double and triple bonds ***are always drawn***.

Example: Propane (C_3H_8)

The **connectivity** of the carbon and hydrogen atoms can be seen ___________ in propane's ***line bond structure***:



In ***condensed structures***, ____________bonds between carbon and hydrogen atoms ***are omitted***. We can draw the condensed structure of propane as:

$CH_3-CH_2-CH_3$ <u>or</u> 

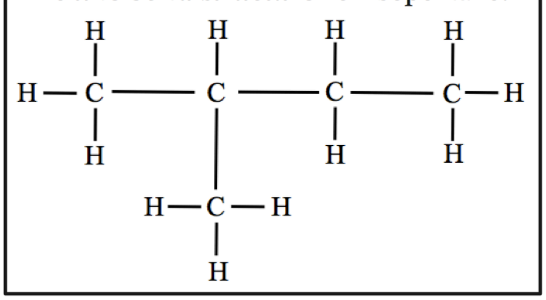
When drawing condensed structures, for ***non*** auto-graded use, carbon-carbon single bonds can be omitted when the carbon atoms are in a ____________ _______________ such as the three carbons in propane:

$$CH_3CH_2CH_3$$

Even though not all of the bonds are drawn in condensed structures, they are ____________and can be *unequivocally* known.

Example: Isopentane **(**C_5H_{12}**)**

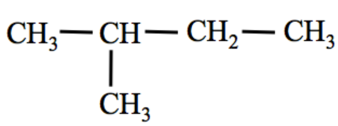
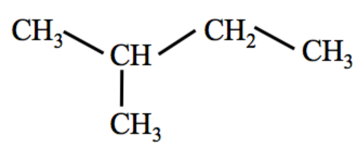
The ***line bond structure*** for isopentane:



In isopentane, there are four carbons bonded in a ***linear sequence***.

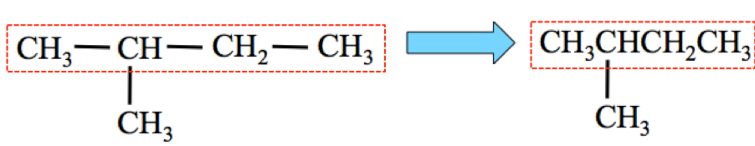
There is one carbon that "________" from the *linear sequence*.

In ***condensed structures***, **single bonds** between carbon and hydrogen atoms ***are omitted***. We can draw the condensed structure of isopentane as:

 <u>or</u> 

When you are asked to draw *condensed structures* for auto-graded, online problem sets you should draw **all of the carbon-carbon bonds** *as shown* ***above***.

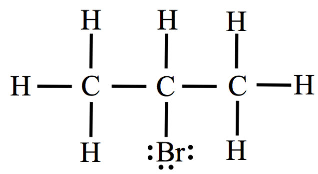
- For **non auto-graded** use, further condensation is possible.
- ***Single bonds*** between **carbon** atoms ***in a linear sequence*** can be omitted as in the case of the four carbons in a **linear sequence** in isopentane (as shown below).



If you **are not** a Saddleback College student, check with your instructor as to which form of condensed structures she/he wishes you to use.

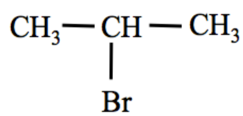
Example: Isopropyl bromide (C_3H_7Br)

The ***line bond structure*** of isopropyl bromide is:

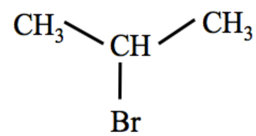


In ***condensed structures***, single bonds between atom pairs ***other than*** **C-C**, **C-H**, **O-H**, and **N-H** are ***always drawn***.

- Therefore the **C-Br** bond is drawn in the condensed structures:

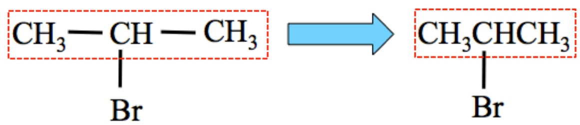


or



- Lone pairs ***can be omitted*** in condensed structures.

- For **non auto-graded** use, further condensation is possible.
- ***Single bonds*** between **carbon** atoms ***in a linear sequence*** can be omitted as in the case of the three carbons in a **linear sequence** in isopropyl bromide.



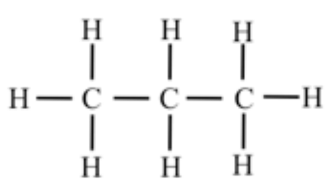
Skeletal Structural Formulas (Skeletal Structures)

Another *structural formula* that chemists use in order to more easily and quickly draw molecules is called the _______________ **structure**.

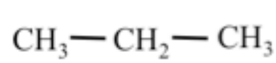
When drawing **skeletal structures**:

- Carbons are not drawn; they are ***implied*** to exist where lines (bonds) _________ or at the ______ of a line (bond).
- Hydrogens are ***omitted*** **if** they are bonded to carbon.
- Bonds between oxygen or nitrogen ***and*** hydrogen may be omitted.
- Atoms other than carbon and hydrogen are ***always*** drawn.
- Lone pairs ***can be*** omitted.

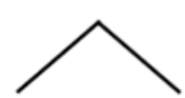
Example: Propane (C_3H_8)



Line Bond Structure



Condensed Structure



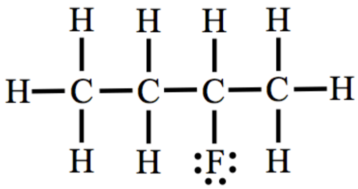
Skeletal Structure

If the hydrogens bonded to carbon are omitted from skeletal structures, how do we determine how many hydrogens are bonded to each carbon? The answer to this is quite simple and is related to the octet rule:

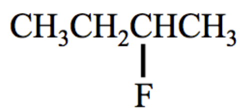
- We ***never*** find ***lone pairs*** on carbon in any molecule that contains more than one carbon atom.

- The octet rule requires that carbons will **always** have ***four pairs of shared electrons around them***, therefore we can deduce the number of *hydrogens* that are bonded to each carbon in a skeletal structure.

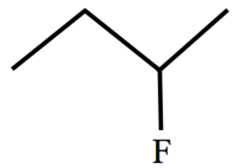
Example: 2-Fluorobutane



Line Bond Structure

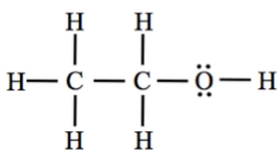


Condensed Structure



Skeletal Structure

Example: Ethanol



Line Bond Structure

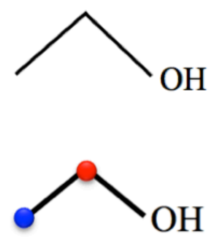
$CH_3—CH_2—OH$

online, auto-graded style

$CH_3CH_2—OH$ **or** CH_3CH_2OH

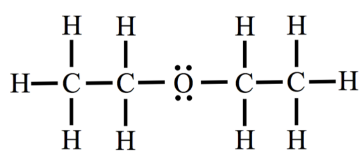
acceptable non auto-graded styles

Condensed Structure



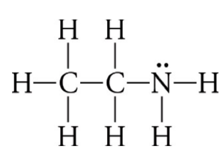
Skeletal Structure

Understanding Check: Draw **condensed *and* skeletal structur** of diethyl ether.



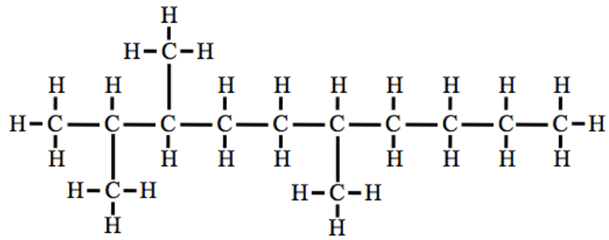
diethyl ether

Understanding Check: Draw **condensed *and* skeletal structure** of ethylamine.



ethylamine

Understanding Check: The line bond structure of a large hydrocarbon molecule is shown below. Draw the **condensed structure** **<u>and</u>** the **skeletal structure** of this molecule.

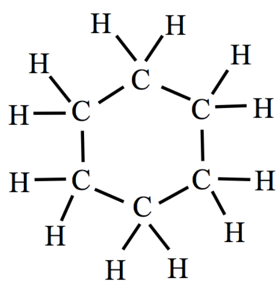


Structural Formulas of Cyclic Compounds

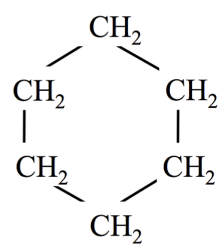
Many compounds contain atoms bonded, not in a *linear sequence*, but in a "**ring**" pattern.

Molecules that contain rings of atoms are called ______________ compounds.

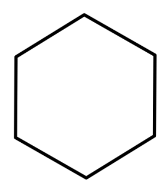
- An example of a *cyclic compound* is a molecule called cyclohexane. The three structural formulas for cyclohexane are:



Line Bond Structure



Condensed Structure



Skeletal Structure

Molecular Geometry

Let's define **molecular geometry** as the ______________ ***arrangement*** of a molecule's atoms.

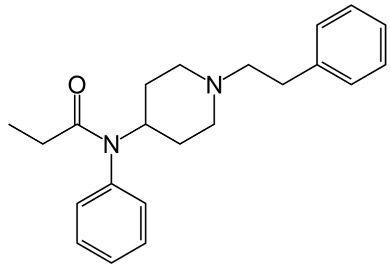
Molecules very rarely take the shapes that are drawn in their structural formulas.

Structural formulas, whether line bond, condensed, or skeletal, are either **one-** or **two-dimensional**.

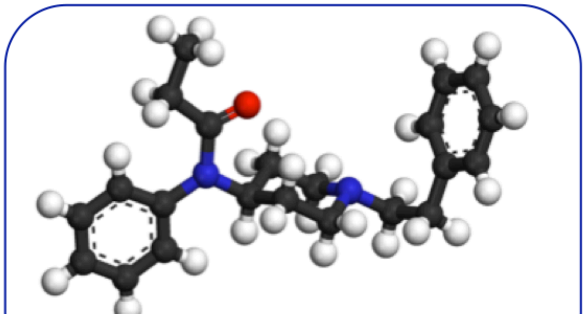
Most molecules are ____________ ***dimensional***.

- Line bond structures are commonly drawn with the lines/bonds at 90^o angles, but the **bond angles** in molecules (the angle between covalent bonds), with rare exceptions, are not at 90^o.
- Skeletal and condensed structures often use bond angles of 90^o or approximately 120^o only for convenience in depicting the bonding pattern connectivity, not necessarily to represent the actual bond angles/geometry of the molecule.

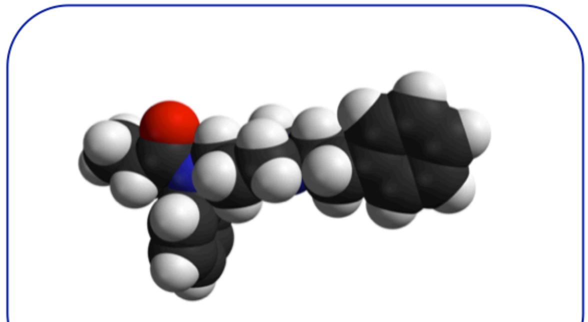
Example: Fentanyl

The ***skeletal structure*** of fentanyl is: 

The "***geometry***" of the molecule can be depicted in two types of views:



ball-and-stick model

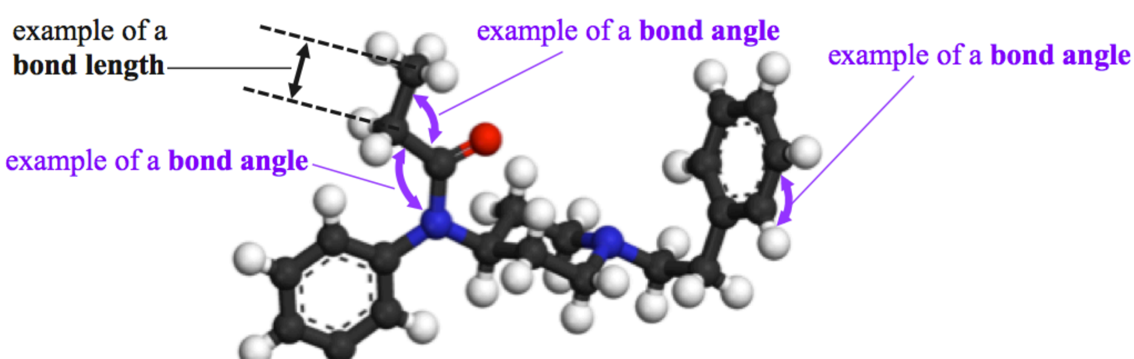


space-filling model

Definition: a _________ ____________ is the angle between two covalent bonds.

Structural formulas, whether line bond, condensed, or skeletal, *do not **explicitly** show the* ________ *bond angles within molecules*. However, **the actual bond angles can be _______________ from a molecule's structural formula**.

To know the ***exact*** molecular geometry, we must know all of the **bond angles** *and* **bond lengths** (distance between bonded atoms) within a molecule.

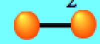
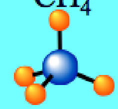
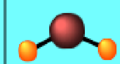
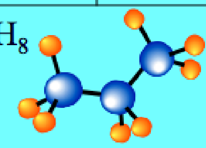


ball-and-stick model of fentanyl

We will **ignore** variations in bond lengths since these relatively minor deviations in length *do not significantly affect the concepts to be discuss*ed.

Therefore, for our purposes, ***predicting molecular geometry will be equivalent to predicting the* bond angles *within a molecule***.

When discussion molecular geometry and other concepts in this course, we will often use the following molecular size categories:

Molecular Size Category	Description	Example(s)	
Diatomic Molecule	Molecule contains **only two atoms**	H_2 	HCl 
Small Molecule	Molecule has **<u>one</u> central atom** with ***<u>all</u> other*** atoms bonded to the central atom	CH_4 	H_2O 
Large Molecule	Larger than *Small Molecule*; there **is <u>not</u> just one single, central atom** with **<u>all</u>** the other atoms bonded to it	C_3H_8 	

For ***diatomic molecules***, the geometry of the molecule is always ______________; the two atoms that make up the molecule exist on the same ***line***.

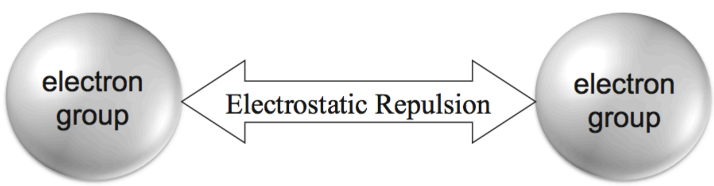
- Examples: H_2 and HCl

For small and large molecules: **bond angles** can be predicted from the _______________ ***formula*** using a model called **Valence Shell Electron Pair Repulsion (VSEPR) Theory.**

Valence Shell Electron Pair Repulsion (VSEPR) Theory: Predicting Bond Angles

The bond angles around *any particular atom of interest* in a molecule can be predicted because the ***groups of electrons*** surrounding this atom ***will*** ___________***each other***.

- These ______________ ______________ (abbreviated as **EG**) consist of, as the name implies, groups of electrons that are localized to a certain area.



You can easily recognize an **EG** by looking for one of the following:

(1) **A** ______________ ____________ (an atom **bonded to** the *atom of interest*).

- Each *bonded atom* counts as **one** electron group.
- Whether the bonded atom is connected to the atom of interest with a **single**, **double**, or ***triple bond***, *all of these shared electrons are localized within a particular region and therefore count as* ***one EG***.

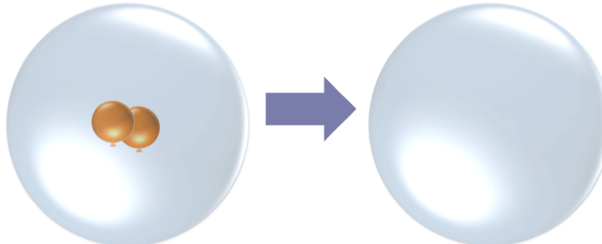
(2) **A** ________ _________ (**on** the *atom of interest*).

- Each lone pair counts as ***one electron group*** (EG
- Even though lone pairs are not attached to other atoms, they do occupy a localized area around the atom of interest and therefore repel other electron groups.

Example: 2 Electron Groups

What do you think will happen to these two balloons if they both have a negative charge?

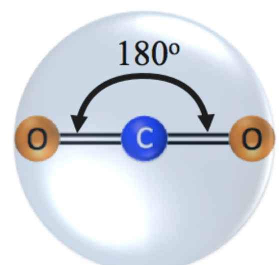
Draw the balloons inside the right-most sphere.



Electrostatic ______________ will cause the balloons to move as far apart from each other as possible.

The same thing happens with _______________ _________________.

Draw the line bond structure for $\mathbf{CO_2}$.



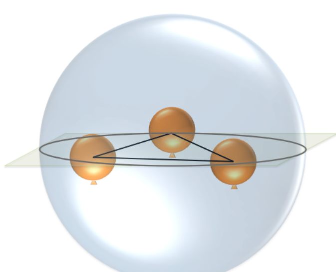
The central atom (C) goes in the ______________ of the sphere.

- There are ________ electron groups around the central atom!
- Each bonded atom counts as ______ electron group.

The electron groups are placed as __________ apart from each other as possible!

Both electron groups are on the same ________ and are at a _________ angle.

Example: 3 Electron Groups



If 3 negatively charged balloons are placed into a hollow, clear plastic sphere, electrostatic repulsion will cause the balloons to move as far apart from each other as possible.

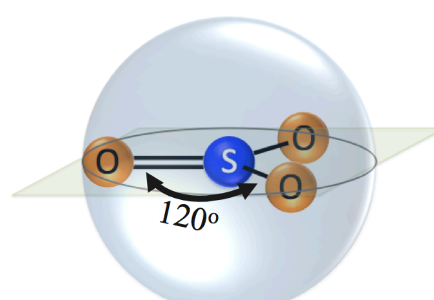
The same thing happens with ***electron groups***.

Draw the line bond structure for $\mathbf{SO_3}$.

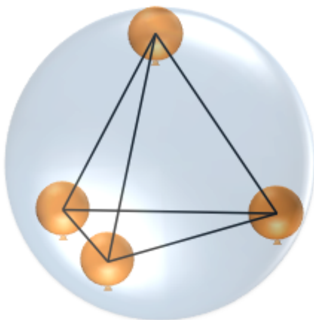
The central atom (S) goes in the middle of the sphere.

There are ________ electron groups around the central atom!

All electron groups are on the same _________ and are at ________ angles.



Example: 4 Electron Groups

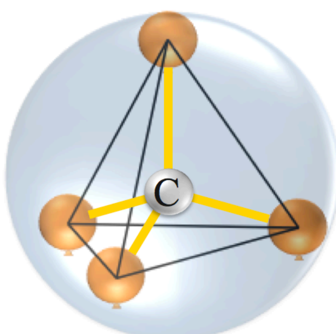


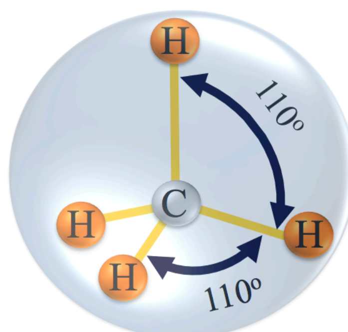
If 4 negatively charged balloons are placed into a hollow, clear plastic sphere, electrostatic repulsion will cause the balloons to move as far apart from each other as possible.

The same thing happens with ***electron groups***.

Draw the line bond structure for a methane molecule (**CH_4**).

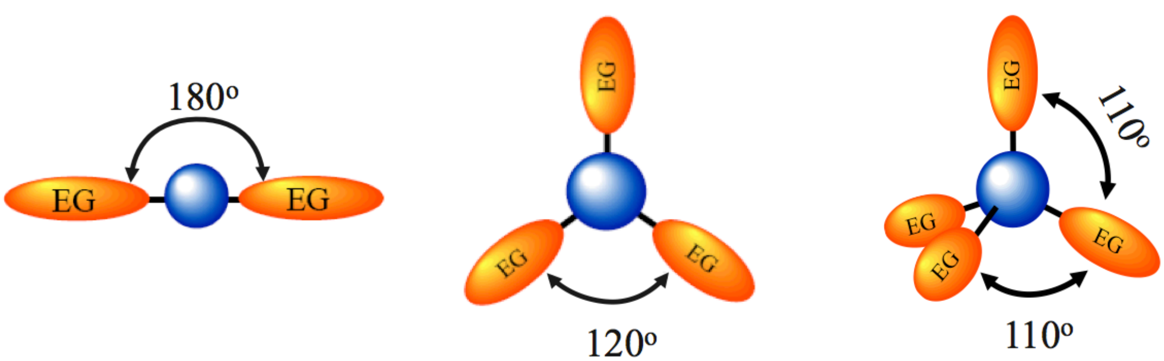
The central atom (C) goes in the middle of the sphere.
How many electron groups are around the central atom?
There are <u>**4**</u> electron groups around the central atom!
The electron groups are placed as far apart from each other as possible!





This arrangement is a 4-sided, 3-dimensional structure with electron groups at about __________ angles.

Electron Arrangement Review

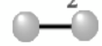
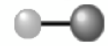


2 Electron Groups 3 Electron Groups 4 Electron Groups

For two or three electron groups, the angles will deviate slightly from **120°** and **110°** when one (or more) of the electron groups is a **lone pair**.

- In this course, these slight deviations in bond angles can be *ignored* because their effects are **not significant** in the chemistry that will be discussed.

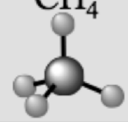
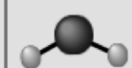
The Geometry of Diatomic Molecules

Molecular Size Category	Description	Example(s)	
Diatomic Molecule	Molecule contains **only two atoms**	H_2 	HCl 

For ***diatomic molecules***, the geometry of the molecule is always **linear**; the two atoms that make up the molecule exist on the same line.

- Since there is only one bond (whether single, double, or triple), it does not make sense to talk about a bond angle.
- To discuss a bond angle, there must be two bonds originating from the same atom.

The Molecular Shape of Small Molecules

Molecular Size Category	Description	Example(s)	
Small Molecule	Molecule has **<u>one</u> central atom** with ***<u>all</u> other*** atoms bonded to the central atom	CH_4 	H_2O 

For ***small molecules***, there exists one **central atom** with all other atoms bonded to the central atom, therefore the *only* bond angles that exist are between bonds __________________ from the **central atom**.

This simplified geometry allows us to categorize (name) particular ***molecular*** ___________ seen in **small molecules**.

Given the line bond structure, small molecules can be categorized by their **molecular shape** as follows:

Step 1. Get the Electron Group Arrangement

Find the ***angle*** between the electron groups and their arrangement around the **central atom**.

- **Two electron groups:** 180^o angle, both electron groups are on the same line as the central atom.
- **Three electron groups:** 120^o angles, all electron groups are on the same plane as the central atom.
  - Angles will deviate **slightly** from 120^o if *one of the electron groups* is a ***lone pair***
- **Four electron groups:** 110^o angles, occupies three dimensions in space.
  - Angles will deviate **slightly** from 110^o if *one or more of the electron groups* is a ***lone pair***

Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.

Small molecules are categorized according to the arrangement of the atoms **only**.

Although **lone pairs** are also electron groups and are therefore important in determining the electron group arrangement in Step 1, they are treated as being "_______________" when assigning molecular shape category names.

Example: The Molecular Shape of CO_2

Draw the line bond structure of CO_2

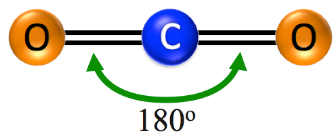
Step 1. Get the Electron Group Arrangement

How many electron groups are around the central atom? ___

What is the bond angle in a 2-electron group arrangement?

Choices: a) 180° b) 120° c) 110°

Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.



This molecular shape is called ______________________.

A very helpful tool for determining **molecular shape** is called the "**ABE Method**."

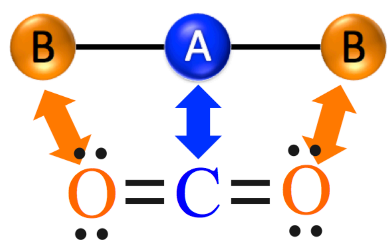
"**A**" represents the central atom (**carbon** in our **CO_2** example).

"**B**" represents atom(s) bonded to the ***central atom***.

(each **oxygen** in our CO_2 example)

"**E**" represents lone pair(s) ***on the central atom***.

(none seen in CO_2 since there are no lone pairs ***on the central atom***)



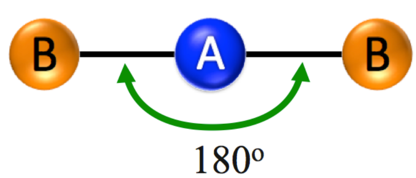
The **ABE** Method uses a "general notation" that indicates the type and number of

_________ __________ (**B** or **E**) that surround the central atom (**A**).

We write the **ABE** Method notation for CO_2 and ***all*** other small molecules that have ***two*** **bonded atoms** and ***no*** lone pairs *on the central atom* as "________."

- There are two oxygens bonded to the central atom therefore we write "**B_2**."

All AB_2 molecules have _______ **bond angles** and are in the _____________ **molecular shape** category.



Molecular Shape Table

Number of Lone Pairs

Number of Electron Groups (EG)

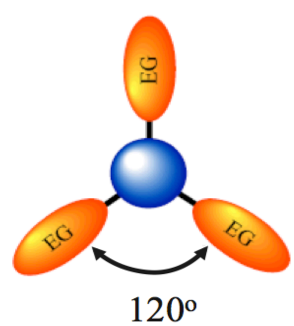
	0 Lone Pairs	1 Lone Pair	2 Lone Pairs
2 EG			
3 EG			
4 EG			

Three Electron Groups

Example: Draw the line bond structure of $\mathbf{SO_3}$

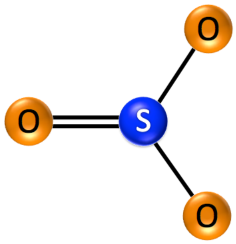
Step 1. Get the Electron Group Arrangement

- The number of electron groups *around the central atom* determines the **bond angles**.

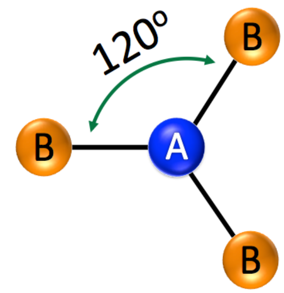


3 Electron Groups

Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.



This molecular shape is called ______________ ___________________.



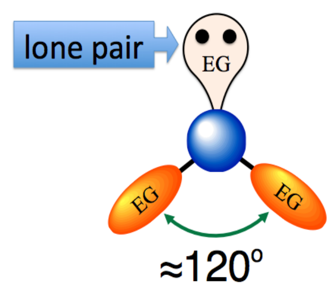
We write the **ABE** Method notation for SO_3 and ***all*** other small molecules that have ***three* bonded atoms** and ***no*** lone pair *on the central atom* as "________."

- There are three oxygens bonded to the central atom therefore we write "**B_3**."

All AB_3 molecules have _______ **bond angles** and are in the ***trigonal planar*** molecular shape category.

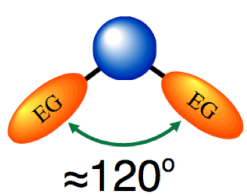
Example: Draw the line bond structure of **ozone gas (O_3).**

Step 1. Get the Electron Group Arrangement



- The number of electron groups *around the central atom* determines the **bond angles**.
  - **Three electron groups:** 120° angles, all electron groups are on the same plane as the central atom. Angles will deviate **slightly** from 120° if *one of the electron groups* is a ***lone pair***.

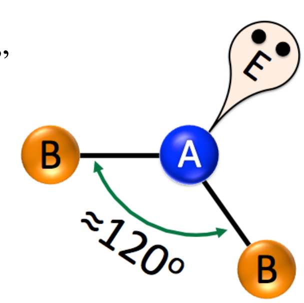
Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.



- When determining **molecular shape**, we treat lone pair(s) as if they were **invisible**!

This molecular shape is called ____________.

We write the **ABE** Method notation for O_3 and ***all*** other small molecules that have ***two* bonded atoms** and ***one*** lone pair *on the central atom* as "__________."



- There are two oxygens bonded to the central atom therefore we write "**B_2**."
- There is **one lone pair** on the *central atom* therefore we write "**E**."

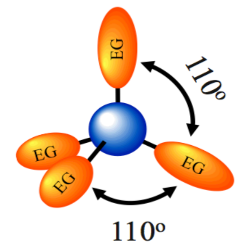
All AB_2E molecules have _________ **bond angles** and are in the ***bent*** molecular shape category.

Four Electron Groups

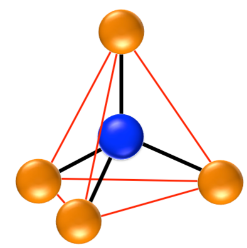
Example: Draw the line bond structure of **methane gas (CH_4).**

Step 1. Get the Electron Group Arrangement

- The number of electron groups *around the central atom* determines the **bond angles**.
  - **Four electron groups:** 110° angles, occupies three dimensions in space

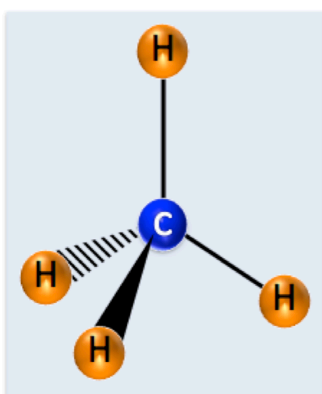


Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.

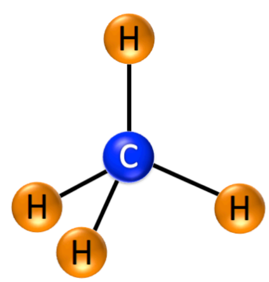


This molecular shape is called ________________________ .

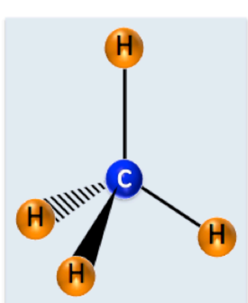
We often use a __________ _____ _________ representation in order to illustrate a three-dimensional molecule on a two-dimensional surface such as a page or computer screen.



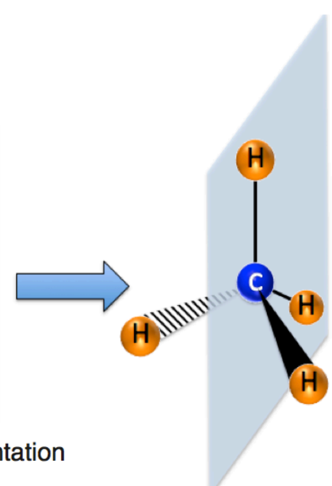
a wedge and dash representation



a perspective representation



a wedge and dash representation

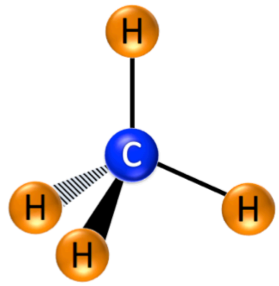


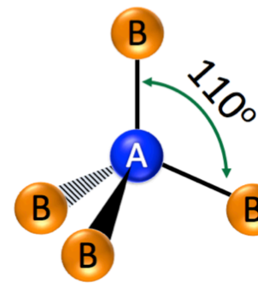
The implication of a wedge and dash representation

Solid wedges indicate bonds that would be coming *out and above* the page (toward the viewer) in the three-dimensional object.

Dashed shapes indicate bonds that would be coming *out and behind* the page (away from the viewer) in the three-dimensional object.

Regular lines (neither wedge nor dash) indicate bonds that exist on the plane of the page in the three-dimensional object.



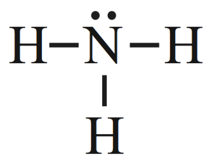


We write the **ABE** Method notation for CH_4 and ***all*** other small molecules that have ***four*** **bonded atoms** and ***no*** lone pair *on the central atom* as "__________."

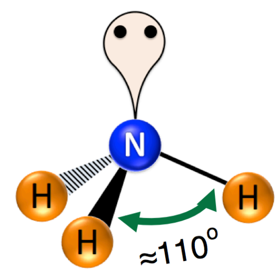
- There are four hydrogens bonded to the central atom therefore we write "**B_4**."

All AB_4 molecules have ________ **bond angles** and are in the ***tetrahedral*** molecular shape category.

Four Electron Group Example: Consider the line bond structure of **ammonia** (**NH_3**).



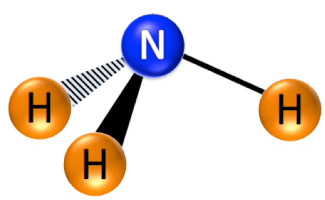
Step 1. Get the Electron Group Arrangement



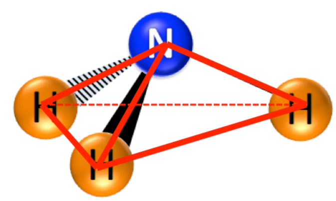
- The number of electron groups *around the central atom* determines the **bond angles**.
  - **Four electron groups:** 110° angles, occupies three dimensions in space. Angles will deviate **slightly** from 110° if *one or more of the electron groups* is a ***lone pair***.

Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.

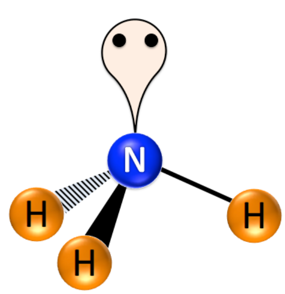
- When determining **molecular shape**, we treat lone pair(s) as if they were **invisible**!

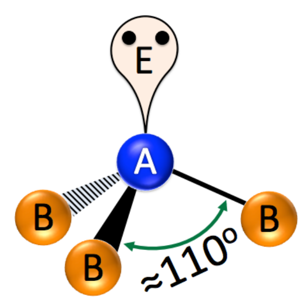


This molecular shape is called _______________.



We write the **ABE** Method notation for NH_3 and ***all*** other small molecules that have ***three*** **bonded atoms** and ***one*** lone pair *on the central atom* as "_________."

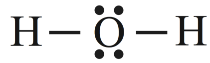




- There are three hydrogens bonded to the central atom therefore we write "**B_3**."
- There is **one lone pair** on the *central atom* therefore we write "**E**."

All AB_3E molecules have ________ **bond angles** and are in the ***pyramidal*** molecular shape category.

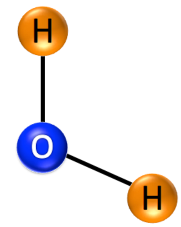
Four Electron Group Example: Consider the line bond structure of a **water molecule (H_2O):**



Step 1. Get the Electron Group Arrangement

- The number of electron groups *around the central atom* determines the **bond angles**.
    - **Four electron groups:** 110° angles, occupies three dimensions in space. Angles will deviate **slightly** from 110° if *one or more of the electron groups* is a ***lone pair***.

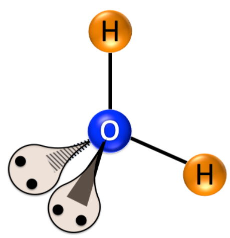
Step 2. Determine the Molecular Shape Category Based on the Arrangement of Atoms.

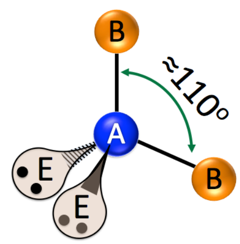


- When determining **molecular shape**, we treat lone pair(s) as if they were **invisible**!

This molecular shape is called _______________.

We write the **ABE** Method notation for H_2O and ***all*** other small molecules that have ***two* bonded atoms** and ***two*** lone pairs *on the central atom* as "_________."

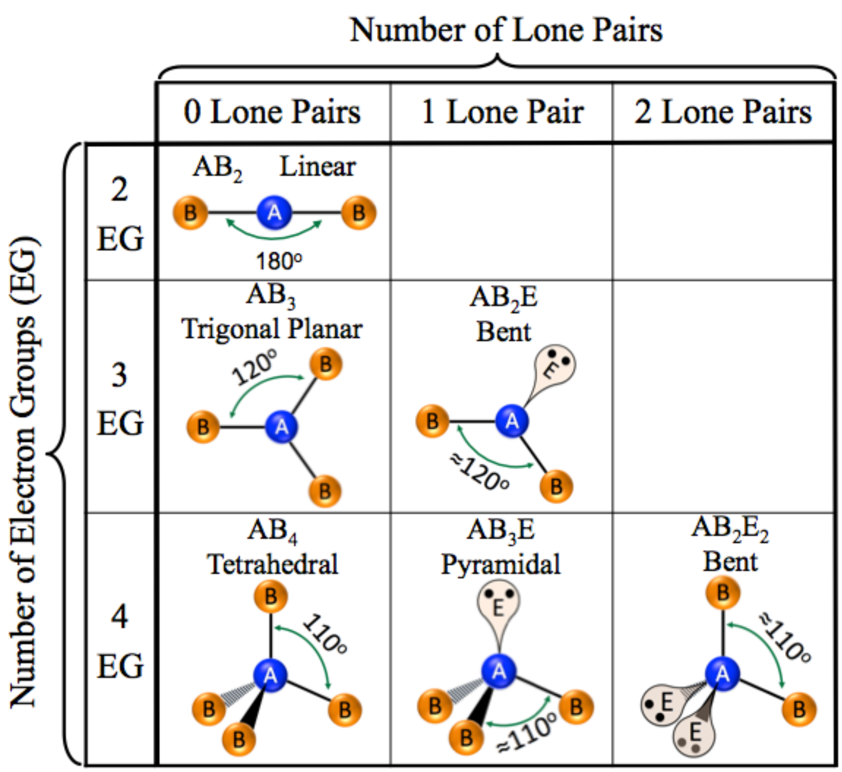




- There are two hydrogens bonded to the central atom therefore we write "$\mathbf{B_2}$."
- There are **two lone pairs** on the *central atom* therefore we write "$\mathbf{E_2}$."

All AB_2E_2 molecules have __________ **bond angles** and are in the ***bent*** molecular shape category.

Molecular Shape Table



Understanding Check: Molecular Shapes for Small Molecules

Determine the **bond angles** (or approximate bond angles) and the **molecular shape** category for each of the following *small molecules*:

a) NF_3

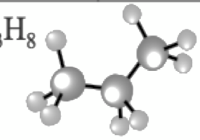
b) H_2S

c) SO_2

d) CS_2

e) carbon tetrachloride

The Molecular Geometry of Large Molecules

Large Molecule	Larger than *Small Molecule*; there **is not just one single, central atom** with **all** the other atoms bonded to it	C_3H_8 

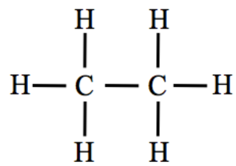
The **bond angles** around any _________ *of interest* in a **large molecule** can be predicted in the same manner that we used for small molecules.

- Simply use the **ABE** method, but in this case let "**A**" represent the ***atom of interest*** in the large molecule instead of the central atom of a small molecule.

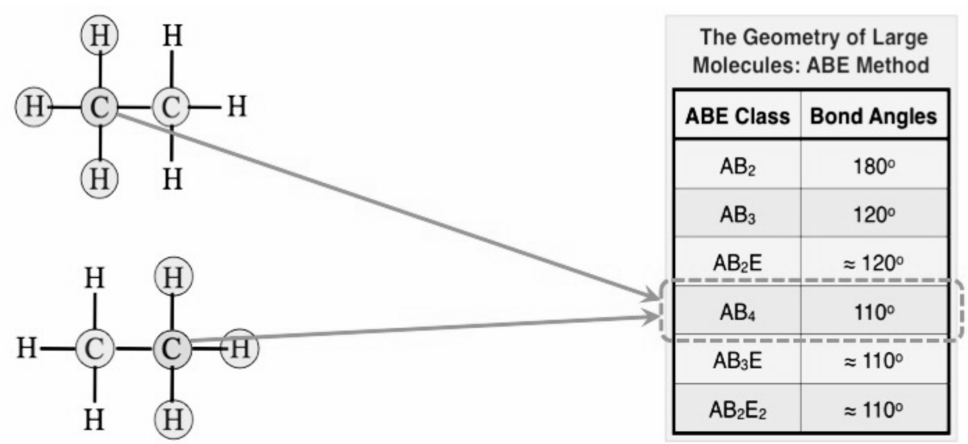
The Geometry of Large Molecules: ABE Method

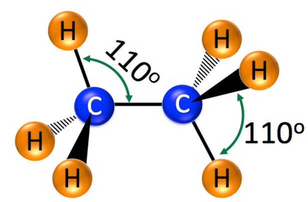
ABE Class	Bond Angles
AB_2	180°
AB_3	120°
AB_2E	≈ 120°
AB_4	110°
AB_3E	≈ 110°
AB_2E_2	≈ 110°

Example: Consider the line bond structure of an ethane molecule:

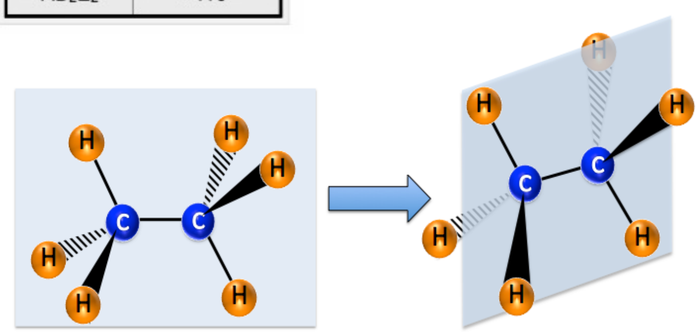


What are the ***bond angles*** around **each** of the **carbon atoms** in an ethane molecule?

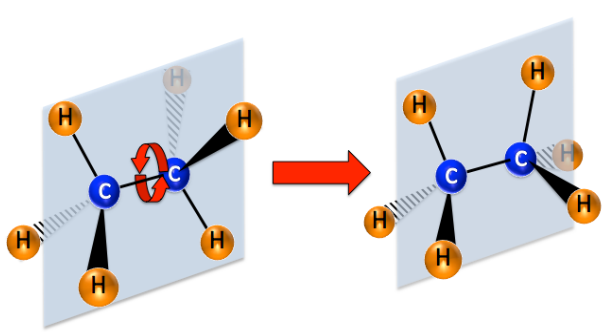




The illustration on the right shows the significance of lines, wedges, and dashes implied in the wedge and dash representation of the ethane molecule.



You may have noticed that we only discussed **molecular geometry (*bond angles)*** for ***large molecules***; we do not categorize large molecules by "*molecular shape*" as we did for small molecules.

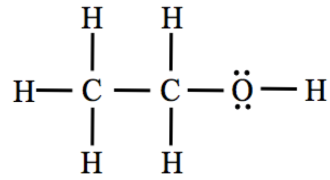


The reason for this is that _______bonds can "_________" and therefore large molecules are **always *changing their shape***.

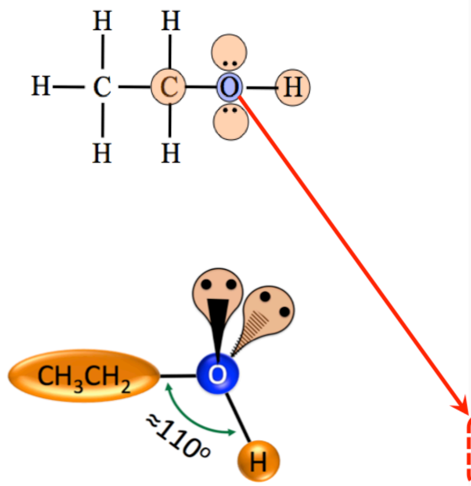
It is important to understand that although single bonds are capable of rotation, ***the bond angles around any particular atom do not change upon rotation***.

Only single bonds freely rotate; because of the nature of ***double*** and ***triple bonds***, they *do not* rotate.

Another example for the geometry of large molecules: Consider the line bond structure of an *ethanol* molecule (shown on the right).



What is the bond angle for the bonds coming from ***the oxygen atom*** in an ethanol molecule?

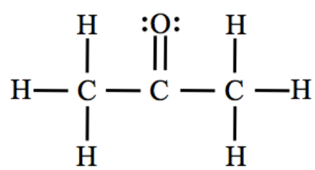


The Geometry of Large Molecules: ABE Method

ABE Class	Bond Angles
AB_2	180°
AB_3	120°
AB_2E	≈ 120°
AB_4	110°
AB_3E	≈ 110°
AB_2E_2	≈ 110°

Understanding Check: Predicting Bond Angles in Large Molecules

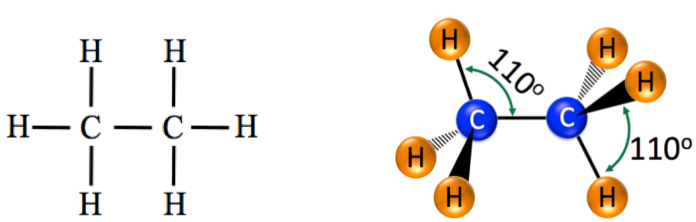
Consider the line bond structure of an acetone molecule:



a) What are the bond angles around the left-most carbon atom?
b) What are the bond angles around the carbon atom in the middle of the molecule?
c) What are the bond angles around the right-most carbon atom?

Reflection and Looking Forward

You have seen how the connectivity of atoms shown in structural formulas, such as line bond structures, can be used to determine molecular geometries (bond angles).
For example:



The molecular geometries are quite important because the observed (macroscopic) properties of a molecular compound are related to its ***nanoscopic*** structure. The *nanoscopic* structure of molecules includes details such as the kind of atoms, the pattern in which they are ***bonded*** to each other, the molecular geometry, and ***how the electrons are distributed within the molecule***.

Throughout the remainder of this course, you will see how the nanoscopic details of a molecule will affect observed, macroscopic properties such as biological effects, melting and boiling temperatures, rates of evaporation, and the ability of one substance to dissolve another substance.

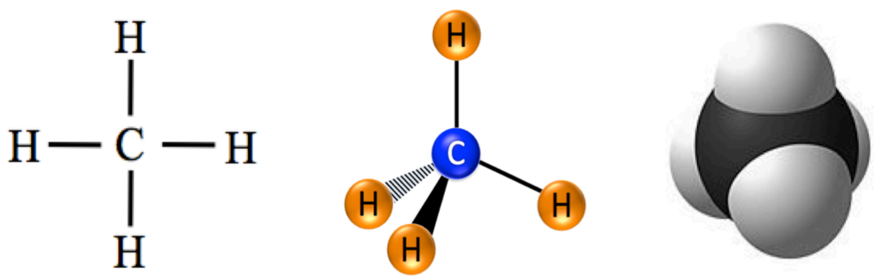
The Distribution of Electrons in Molecules

Molecular geometry plays a major role in ***how the electrons are distributed in molecules***.

The **electron distribution** is a major factor that determines the chemical and physical _______________ of molecular compounds.

- For example, consider two molecules that are quite similar in size - CO_2 and H_2O. Both molecules contain three atoms, however at room temperature they are in different phases; CO_2 exists as a gas, whereas H_2O exists as a liquid (water).

Representations of Methane (CH_4)



line bond structure — ball-and-stick model — space-filling model

Space-filling models are images that show *molecular geometry* and also indicate the approximate *area that electrons occupy.*

Here is a new concept: ***the electrons are*** ____ ***within molecules.***

- Sometimes electrons spend ____________ time in one particular region of the molecule giving that region ***additional*** ________________***charge***.
- Sometimes a particular region of a molecule is ***electron*** ____________ , giving that region of the molecule ***additional*** _______________ ***charge***.

The existence of this ***uneven charge*** within a molecule is a major factor that makes the **properties** of molecules such as CH_4, CO_2, and H_2O, quite different.

Let's first consider the H_2 molecule: H – H

The two electrons in this ***covalent bond*** are "*shared evenly*" by the hydrogens.

The electrons do not spend more time, *on average*, closer to any one of the two hydrogen atoms.

This is because the bond is between ___________ ***atoms***; the electrons are equally attracted to **each** of the positively-charged hydrogen nuclei.

When electrons are shared between ***like atoms***, such as in H_2, they are, on average, **evenly** distributed between the atoms.

Next, let's consider the HCl molecule:

The two electrons in the **H-Cl** ***bond*** are ***not*** shared ***equally*** by the chlorine and hydrogen atoms.

The shared electrons spend, on average, a bit more time nearer the chlorine than the hydrogen.

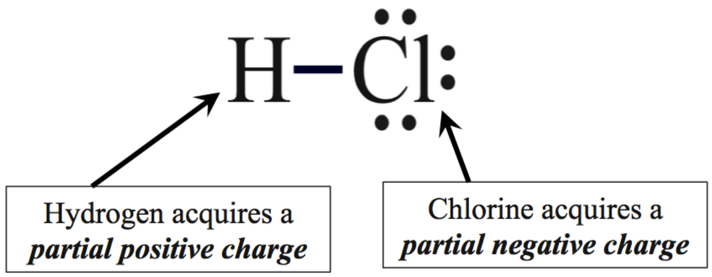
This is because the bond is between _____________ atoms (chlorine vs. hydrogen).

- The shared electrons happen to be more strongly attracted to the chlorine than they are to the hydrogen.

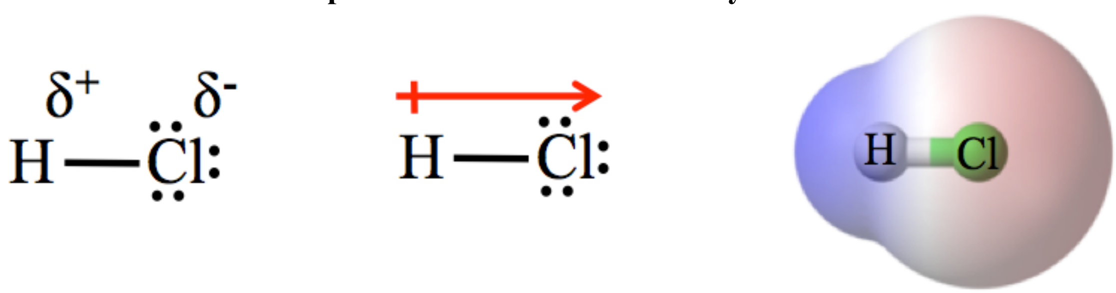
When electrons in bonds are **evenly** shared, as is the case for ***like** atoms* (such as H_2*)*, we call the bond a _____________ _________.

When electrons in bonds are not equally shared, as is the case for *un-like atoms*, we call the bond a _____________ _________.

Scientists and engineers use the term _____________ ___________ to describe this physical state that results when there is a separation of charge over a short distance such as the length of a chemical bond.



Representations of Bond Polarity in HCl



The line bond structure (above, left) uses the Greek lowercase letter delta (**δ**) to indicate a *partial positive charge* (**δ**$^+$) or a *partial negative charge* (**δ**$^-$) near the appropriate atoms.

Another way to show bond polarity and partial charge is to use the **dipole arrow** (above, middle).

- When using *dipole arrows*, the arrow points to the partial negative charged region and a crossed line (looks like + on the end of the arrow) is used to indicate the partial positive charged region.

The space-filling model (above, right) not only shows the region that the molecule's electrons occupy, but is also shaded to indicate the *electron distribution*.

- The area where electrons spend more time and therefore has a partial negative charge is shaded red.
- The area that is electron deficient has a partial positive charge and is shaded blue.

Electronegativity

The relative ability of an atom in a bond to attract electrons is called its _______________________.

- The *more* electronegative atom in the bonded pair will have a stronger attraction to the shared electrons, the electrons will spend a bit more time in the vicinity of that atom, therefore the atom will have a **partial negative charge**.
- The *less* electronegative atom in the bond will have a **partial positive charge**.

Electronegativity Values

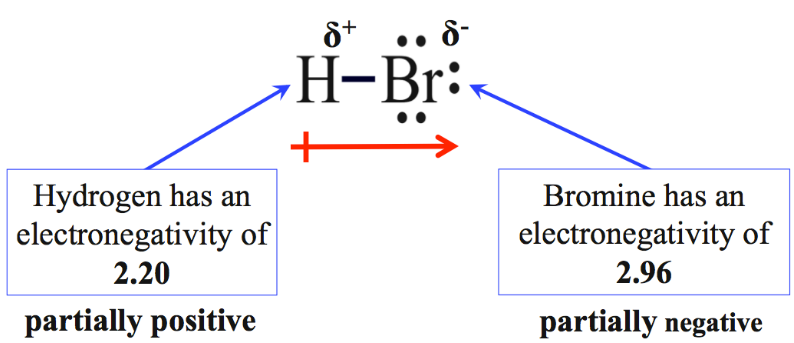
H 2.20																	He
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 1.87	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo

* Lanthanoids	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27
** Actinoids	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3

Note that there is a **periodic trend** in electronegativity values of the **s-** and **p-**block elements; electronegativity increases going from left to right (in a row) and bottom to top (in a column).

You can determine which atom in a polar bond is ***partially positive*** and which is ***partially negative*** by comparing the **electronegativity values** of the two atoms.

Example: Consider the line bond structure of hydrogen monobromide (HBr).



Understanding Check: Bromine monofluoride (BrF) exists as a brown gas and is used as an algaecide, fungicide, and disinfectant in some industrial applications. Draw the line bond structure of bromine monofluoride.

a) Predict which atom is **partially positive** and which atom is **partially negative**.
b) Add the δ^+ or δ^- symbols to the appropriate atom.
c) Add a dipole arrow below the line bond structure to indicate the polarity.

The **degree of bond polarization** depends on the _______________ in electronegativities between the two bonded atoms.

- Greater differences in the abilities of the bonded atoms to attract the shared electrons results in ***more polarized bonds***.

Example: The electronegativity of hydrogen, fluorine, and bromine are 2.20, 3.98, 2.96, respectively. Which bond is ***more polarized***, the **H-F** or **H-Br** bond?

H-F: The ***difference*** in electronegativity between F and H: 3.98 - 2.20 = **1.78**

H-Br: The ***difference*** in electronegativity between Br and H: 2.96 - 2.20 = **0.76**

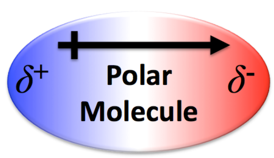
Since there is a greater difference in electronegativity between F and H vs. Br and H, **the H-F bond is __________ polarized than the H-Br bond.**

Summary of Bond Polarity

We can classify covalent bonds as being *either* ___________ or ________________.

- **Polar bonds** occur because of ***unequal sharing*** of electrons in covalent bonds when two ***un-like* atoms** (such as H-Br or H-Cl) are bonded together.
- **Nonpolar bonds** occur when the electrons are shared evenly between two ***like* atoms** (such as H-H or F-F).

The Polarity of Molecules



Similar to the concept of *bond polarity*, when electrons are not symmetrically distributed **in a molecule**, it results in a **polar** _________________; a molecule with one end that has a partial negative charge and one end that has a partial positive charge.

The existence and degree of polarity (strength of the dipole) in molecules has implications in many of the chemistry concepts that you will learn throughout this course.

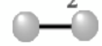
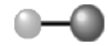
For example, one type of force responsible for attracting molecules ***to other molecules*** that is, in part, responsible for whether the molecules exist close to each other in the liquid or solid phase instead of the gas phase is called the **dipole-dipole force**.

- The partially positive (**δ^+**) end of one molecule's dipole is attracted to the partially negative (**δ^-**) end of another molecule's dipole (and vice versa) by an electrostatic attraction as illustrated below.

In this video, you will learn how to determine if a **molecule** is ***polar*** or ***nonpolar*** based on ***both*** the existence (or not) of _________ _________ **and** its molecular _________________.

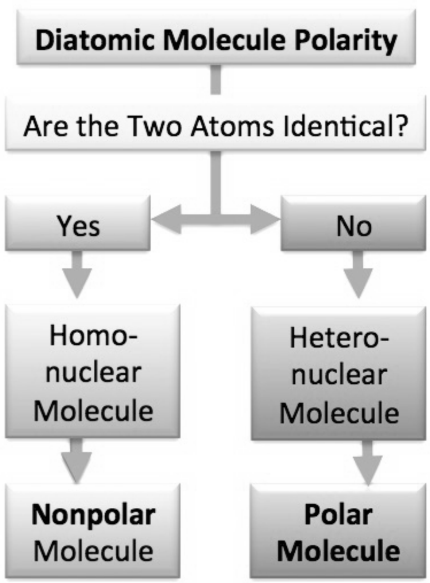
I will show you a method for determining the polarity of a molecules for each of our *molecular size categories* (diatomic, small, and large molecules).

The Polarity of *Diatomic* Molecules

Molecular Size Category	Description	Example(s)	
Diatomic Molecule	Molecule contains **only two atoms**	H_2 	HCl 

Since diatomic molecules contain ***only one bond***, the polarity of the molecule depends only on the polarity of that bond.

- A ___________________ ***diatomic molecule***, such as H_2, contains ***like*-atoms** and a *nonpolar* bond, therefore the molecule is _______________.
- A ___________________***diatomic molecule***, such as HCl, contains two ***un-like* atoms** and a *polar* bond, therefore the molecule is ____________.



Understanding Check:
Determining the Polarity of Diatomic Molecules

Categorize each of the diatomic molecules listed below as **polar** or **nonpolar**.

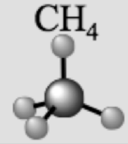
a) O_2

b) carbon monoxide (CO)

c) N_2

d) nitrogen monoxide (NO)

The Polarity of *Small* Molecules

Molecular Size Category	Description	Example(s)	
Small Molecule	Molecule has **<u>one</u> central atom** with ***<u>all</u> other*** atoms bonded to the central atom	CH_4 	H_2O 

Small molecules contain more than one bond.

The *individual bonds within a molecule* can be polar and/or nonpolar.

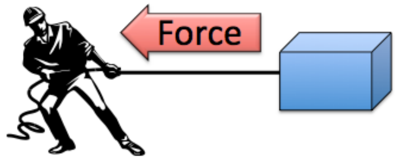
Just because a small molecule contains polar bonds does not necessarily mean that it is a polar molecule.

In order to determine if a small molecule is polar, we must look at what happens when we ***<u>add</u> up*** all of the dipoles.

Dipoles are physical quantities that not only have a magnitude (amount of charge), but also have a direction.

> An analogous physical quantity that you are familiar with is *velocity*. When you drive your car, to completely describe your path you would use both the *speed* (miles per hour) and the *direction* you are traveling.

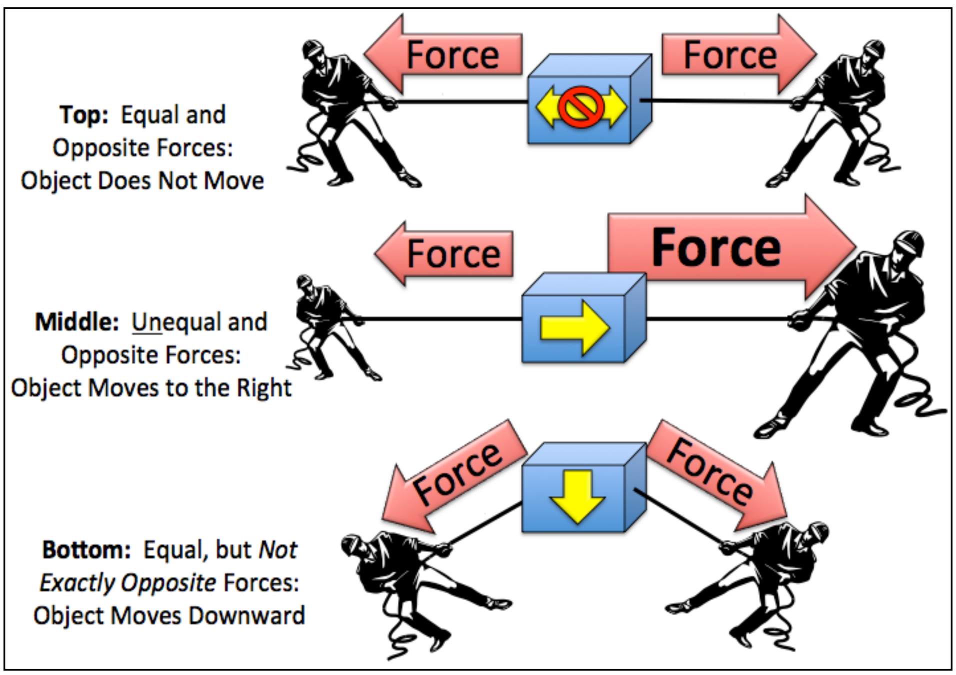
Another physical quantity with both magnitude and direction is a **force**. When you pull on an object, how fast and where the object moves depends on how hard you pull ***and*** in what *direction* you pull.



Quantities with both magnitude and direction are called **vector quantities**.

Covalent bond electric dipoles are vector quantities; they have both *magnitude* (based on differences in electronegativity of the bonded atoms) and *direction* (based on bond angles).

Before we discuss "*adding up*" dipole vectors, let's *add up* some vector quantities of which you have prior knowledge. We will consider adding a couple of force vectors:



Top: Two equal tensions (represented by the red vector arrows of the same length) are exerted in opposite directions. The blue box does not move because the forces are "balanced" or "cancel"; they *add up to zero*.

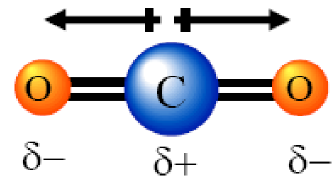
Middle: The two tension forces are in opposite directions, however the force pulling to the right is larger (represented by a larger vector arrow). The blue box moves to the right because the forces are not "balanced" and do not "cancel" - they ***do not*** *add up to zero*.

Bottom: The two tension forces are equal but they are *not oriented in exactly opposite directions.* The blue box moves downward because the forces are not "balanced" and do not "cancel" - they ***do not*** *add up to zero*.

To determine if a *small* molecule is polar:

We must look at what happens when we _____________ up all the dipoles.

Example: CO_2



There are 2 polar bonds in a CO_2 molecule.

This results in two_________________.

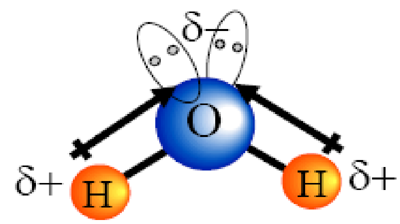
- The oxygens have a _______________ negative charge.
- The carbon has a partial positive charge.

Since the molecule is ___________, the dipoles cancel each other and the molecule is _____________.

- All *symmetric* small molecules are ***nonpolar***.

Let's look at what happens when we add up all the dipoles in another molecule:

Example: H_2O

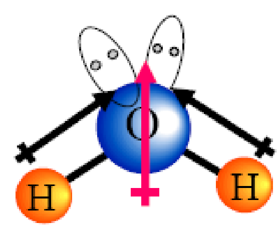


There are 2 polar bonds in an H_2O molecule.

This results in two dipoles.

- The oxygen has a partially negative charge.
- The hydrogens have a partially positive charge.

Since the molecule is <u>not</u> symmetric, the dipoles *do not cancel* each other and the molecule is _____________.



Since dipoles in ***nonsymmetric*** molecules do not cancel each other out, they ***add up*** to yield a **molecule with a dipole**.

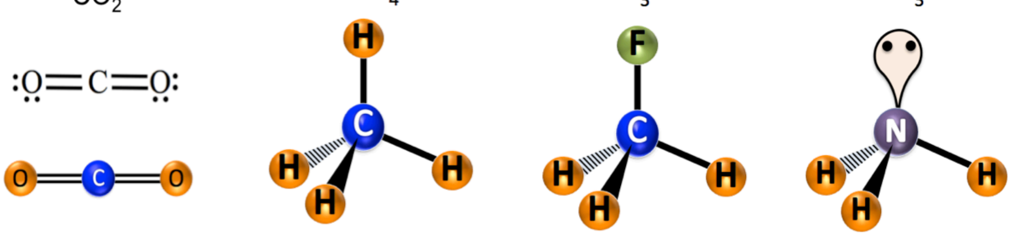
Nonsymmetric small molecules with polar bonds are **polar** molecules.

General Rule to Know if a molecule is Symmetric or Nonsymmetric:

Symmetric molecules have the *central atom* surrounded by _______________________electron groups.

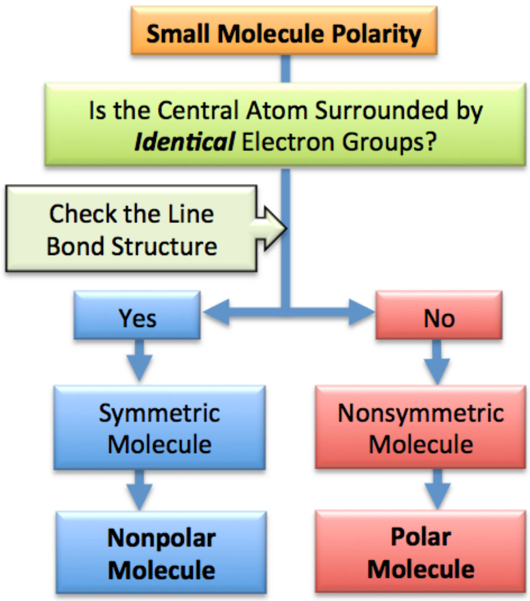
Examples of symmetric and *non*symmetric molecules:

CO_2 | CH_4 | CH_3F | NH_3



Symmetric (Nonpolar) | Symmetric (Nonpolar) | ***Non***symmetric (Polar) | ***Non***symmetric (Polar)

Polarity of Small Molecules Summary



Understanding Check:
Determining the Polarity of Small Molecules

Categorize each of the diatomic molecules listed below as **polar or nonpolar**. Note that you will need to begin with the line bond structure of a molecule in order to determine its polarity.

a) NF_3

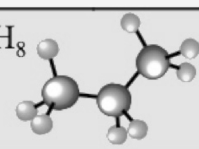
b) H_2S

c) O_3

d) CS_2

e) CCl_4

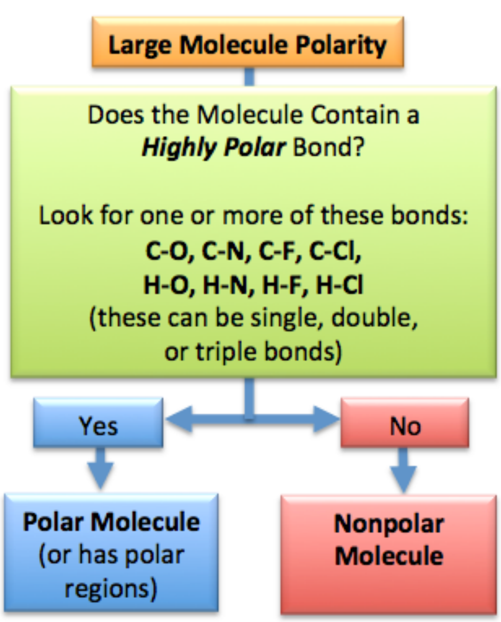
The Polarity of *Large* Molecules

Molecular Size Category	Description	Example(s)
Large Molecule	Larger than *Small Molecule*; there **is not just one single, central atom** with **all** the other atoms bonded to it	C_3H_8 

In this course, we will be working with large molecules (organic and biochemistry).

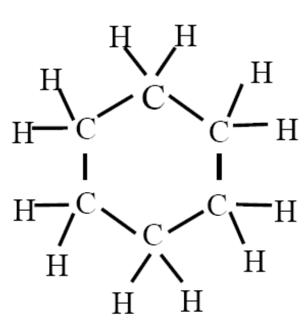
Large molecules will be considered polar (or have polar regions) if they have __________ ________ bonds.

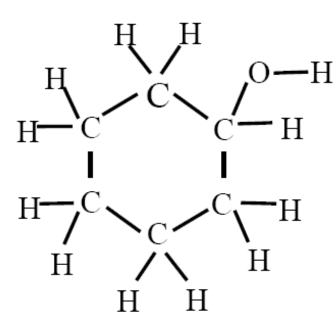
- The ***important highly polar covalent bonds***, especially in organic and biochemistry, are those in which either **hydrogen** or **carbon** atoms are covalently attached to **nitrogen**, **oxygen**, **fluorine**, or **chlorine** atoms.



Example:

Cyclohexane **vs.** **Cyclohexanol**





No *highly polar* bonds

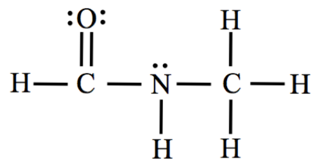
Nonpolar

Two *highly polar* bonds (C-O and O-H)

Polar

Understanding Check: Determining the Polarity of Large Molecules

The compound shown below has been used to kill any insect larvae present in cereal and dried fruit.



This molecule is ________.

a) polar

b) nonpolar

Formal Charge

There is ***one more*** phenomenon that results in the *uneven distribution of charge*.

We will use the concept of "_____________ ______________" to account for electrically-charged regions within molecules (and polyatomic ions).

The ***formal charge*** concept involves identifying _________ in molecules (or in polyatomic ions) that have an ***excess*** *or* ***deficiency of electrons***.

- An excess of electrons on a particular atom means that there are more electrons in the vicinity of an atom than there are protons in the atom. If there is one extra electron in the vicinity of a particular atom, the *formal charge* of the atom would be 1-.
- Conversely, when we say that an atom is electron deficient, this means that there are fewer electrons in the vicinity of a particular atom than there are protons in the atom. If an atom has a deficiency of one electron (one more proton than electrons), the formal charge of the atom would be 1+.

What Do You Need to Know About Formal Charge?

We will soon be focusing on organic biomolecules. Biomolecules consist primarily of C, H, O, and N.

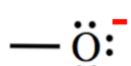
- In the compounds that you will come across in this course, *carbon* and *hydrogen* atoms *will never* have a formal charge.
- *Oxygen* and *nitrogen* ***may or may not*** have a formal charge.

I will only need you to be able to determine the formal charge on two elements -

_____________ and _____________.

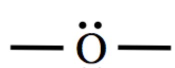
1) The Formal Charge on Oxygen:

- Oxygen has a formal charge of **1-** when it has ***just*** ___________ ***single bond*** (three lone pairs).



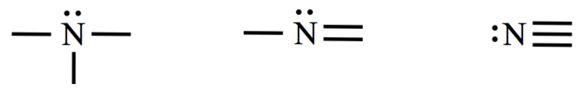
- In ***all other cases that we discuss in this course***, oxygen will **not** have a formal charge (formal charge = zero).

  - Examples:

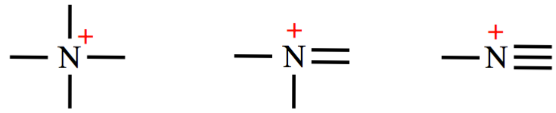


2) The Formal Charge on Nitrogen:

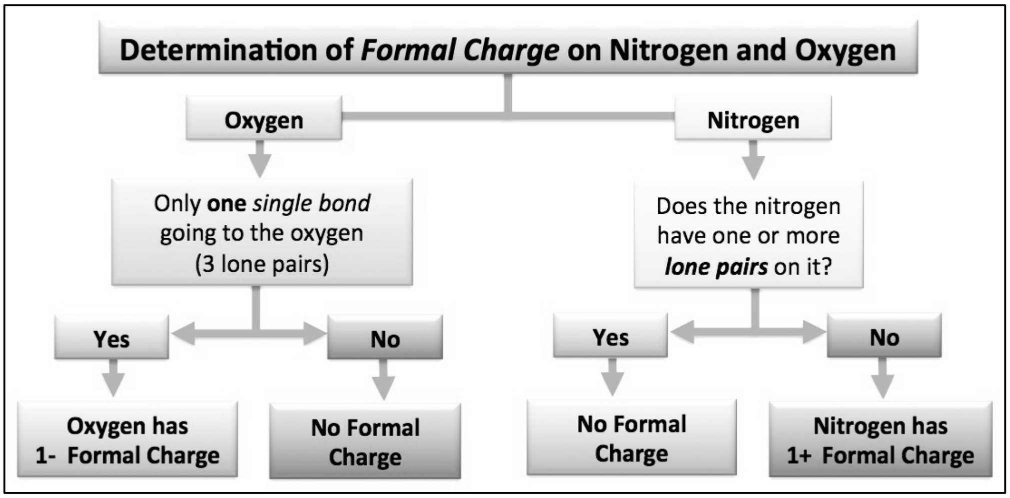
- Nitrogen will have ***no*** *formal charge* when it has ***one or more lone pairs on it***.



- In _______ ***other cases***, nitrogen has a **1+** formal charge.

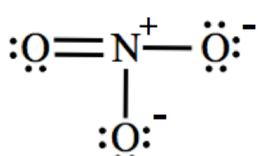


Determination of the formal charge of oxygen and nitrogen is summarized in the thought-map flowchart shown below.



OPTIONAL: If you are interested in learning the details of how formal charges are assigned to oxygen, nitrogen, and other atoms based on their bonding patterns, you can find that information in Appendix 3 of your textbook.

Example: The line bond structure of the nitrate ion (NO_3^-) is shown below. Add the ***formal charge*** (as a superscript) to the line bond structure next to any **oxygen** or **nitrogen** atom that has a non zero formal charge.

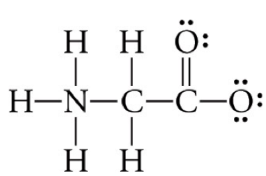


The nitrogen has ***no lone pairs***, therefore it has a formal charge of **1+** (or +).

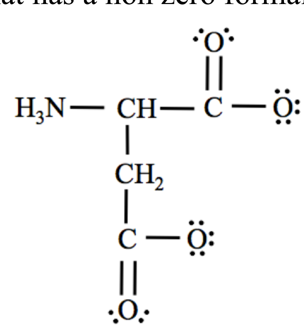
The oxygens below and to the right of the nitrogen have ***just one single bond***, therefore they have a formal charge of **1-** (or **-**).

The oxygen to the left of the nitrogen ***does not have just one single bond*** (it has a *double bond*), therefore it has no formal charge (formal charge = zero).

Understanding Check: The line bond structure glycine, an amino acid present in many proteins, is shown below. Assign formal charges to the nitrogen and each oxygen atom in this compound.



Understanding Check: Aspartic acid is one of the 20 common amino acids that make up the proteins in our bodies. The condensed structure of the aspartic acid is shown below. Add the formal charge (as a superscript) next to any ***oxygen or nitrogen atom*** in the condensed structure that has a non zero formal charge.



Summary

In the last two videos, you learned *two ways* to predict the ***distribution of charge***:

> **Molecular Polarity:** Because of the unequal sharing of electrons, regions of a molecule can have **partial charge**. These charged regions have a *partial positive charge* (**δ^+**) or *partial negative charge* (**δ^-**).
>
> **Formal Charge:** A region in the vicinity of a *particular atom* in a compound can have ***formal charge*** *because of an excess or deficiency* of electron(s) at that atom.

Both the polarity and presence of formal charges in molecules are important in understanding how ***compounds interact with each other***, and also in understanding the forces within large biomolecules such as DNA, RNA, proteins, and carbohydrates. These forces are crucial for the biological molecules maintaining their shape and biological function.

Noncovalent Interactions

Ionic compounds exist as crystalline solids at room temperature because the ions are held together by **ionic chemical bonds**.

Molecular compounds can also exist as crystalline solids.

- Example: H_2O (ice or snow flakes)

H_2O molecules are not ions, so ***what forces are responsible for holding H_2O molecules, or other molecules, together in their solid states****?*

The answer is that there are *electrostatic attractive forces* called ________________ _______________ that attract molecules to other molecules.

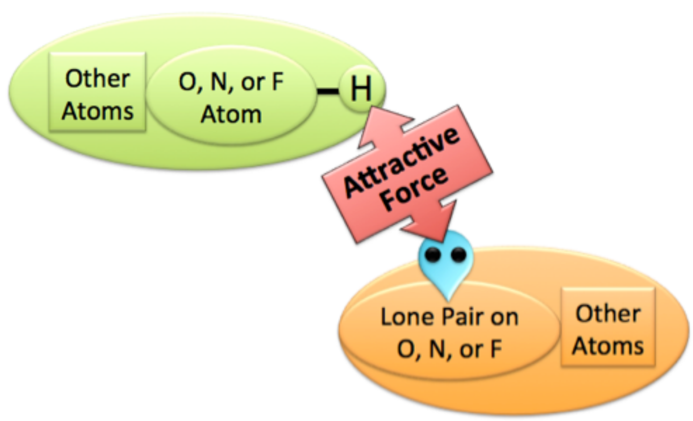
Noncovalent interactions are quite different from the ***covalent bonding forces*** that are present within molecules and polyatomic ions.

- ______________ ***bonding forces*** result from **shared electrons**.
- ***Noncovalent interactions***, as the name implies, ***do not*** involve the sharing of electrons; they are strictly **electrostatic attractions**.

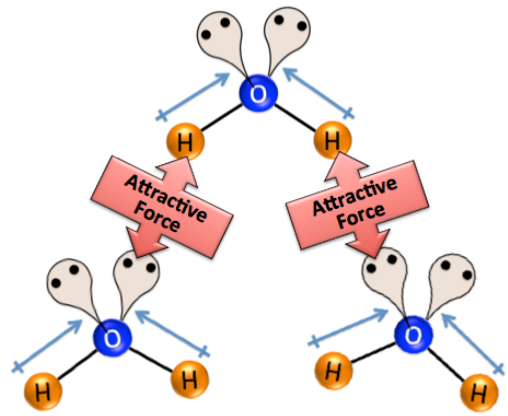
In this video, you will learn about **five** types of ***noncovalent interactions***.

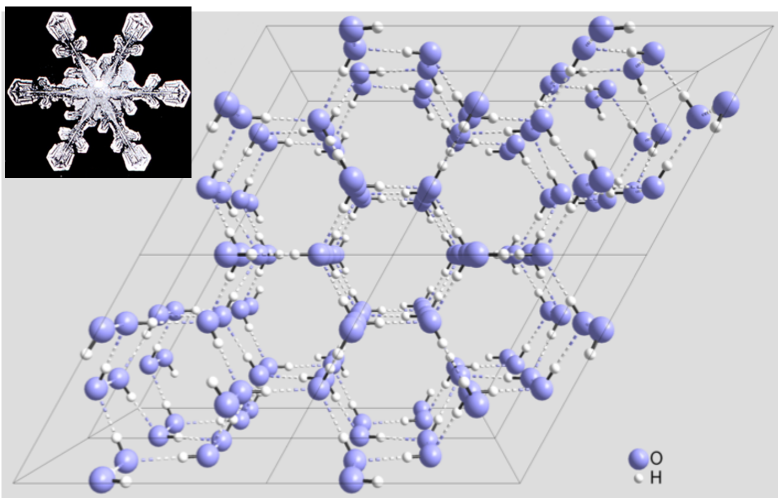
1) Hydrogen Bonding

Hydrogen bonding is the result of the electrostatic attraction *between* the partial positive charged end of a ***particularly strong polar bond*** (O-H, N-H, or F-H) and the negative charge of a lone pair of electrons on a very electronegative atom (O, N, or F).



Example: *Hydrogen bonding* is the major force responsible for H_2O, a relatively small and light molecule, existing in the liquid phase instead of the gaseous phase at room temperature.

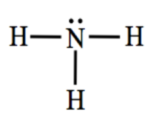




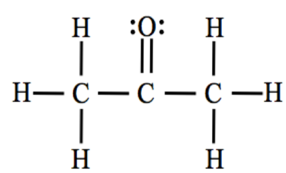
On the Left: Hydrogen bonds (as dashed lines) in a ball and stick model of an ice crystal.

Understanding Check:

The line bond structures of ammonia and acetone are shown below:



ammonia

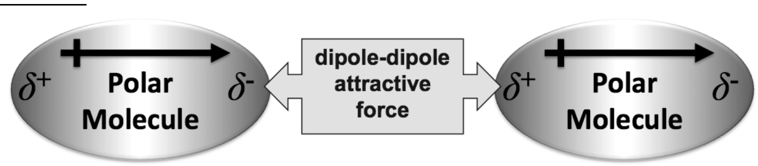


acetone

a. Can hydrogen bonding occur between two ammonia molecules?
b. Can hydrogen bonding occur between two acetone molecules?
c. Can hydrogen bonding occur between an ammonia molecule and an acetone molecule?

2) Dipole-Dipole Forces

Polar molecules are attracted to other polar molecules by a type of noncovalent interaction called the _________________________ **force**.

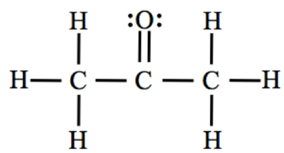


The partially positive (**δ^+**) end of one molecule's dipole is electrostatically attracted to the partially negative (**δ^-**) end of another molecule's dipole (and vice versa).

Example: Diethyl Ether

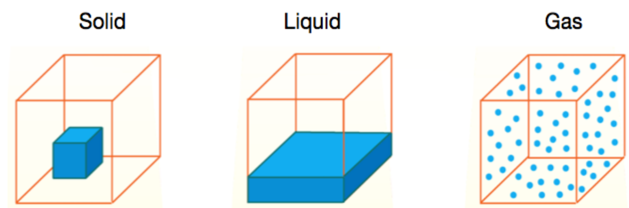
Understanding Check:

a. Can a dipole-dipole force occur between two H_2O molecules?
b. Can a dipole-dipole force occur between two methane (CH_4) molecules?
c. Can a dipole-dipole force occur between two acetone molecules?



acetone

It is ***noncovalent interactions*** that hold molecules together such that they exist in the solid or liquid phases instead of the gaseous phase.



Whether a substance exists in the gas, liquid, or solid phase is determined by a ______________ between noncovalent interactions (working to keep the particles close to one another) and temperature (kinetic energy working to distribute the particles randomly in their container).

- If the noncovalent interactions are dominant, then the substance will exist in the solid or liquid phase.
- If the temperature/kinetic energy can overcome the noncovalent interactions, then the substance will exist in the gaseous phase.

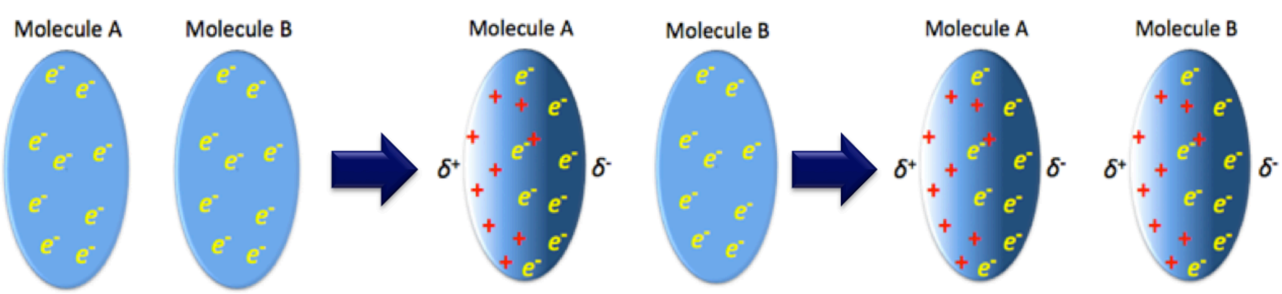
It was observed that nonpolar molecules such as CO_2, O_2, and N_2, which are ***not capable*** of interacting through dipole-dipole forces or hydrogen bonding, can be cooled to temperatures at which they exist in the liquid and solid phases.

In 1930, Fritz London, along with R. Eisenschitz, came up with a model to explain this observation.

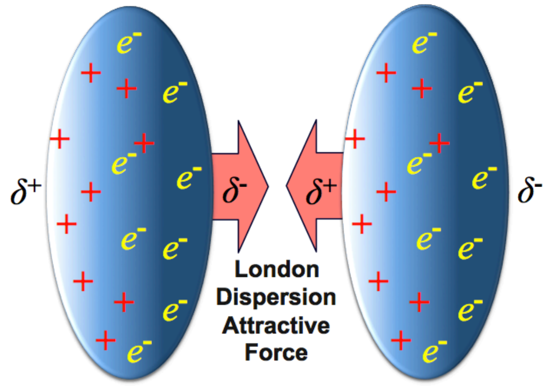
3) **London Dispersion Forces**

The explanation is based on our third type of ***noncovalent interaction***, which came to be known as **London dispersion forces**, named in honor of Fritz London.

London dispersion forces are caused by an "instantaneous" dipole in one molecule _____________ the formation of a "temporary" dipole in another molecule.



_______ ***molecules***, polar and nonpolar, contain electrons and will therefore be attracted to each other through London dispersion forces.



The ___________ a molecule is, the easier (lower in energy) it is to polarize its electrons, and therefore, **the stronger is its London dispersion force interactions**.

- This trend can be observed by noting the boiling points of the molecules in the table below:

Effect of Molecular Size on London Dispersion Forces			
Molecule Name	**Condensed Structure**	**Boiling Point (°C)**	**Phase at Room Temperature**
methane	CH_4	-164	gas
ethane	CH_3CH_3	-89	gas
propane	$CH_3CH_2CH_3$	-42	gas
butane	$CH_3CH_2CH_2CH_3$	0	gas
pentane	$CH_3CH_2CH_2CH_2CH_3$	36	liquid

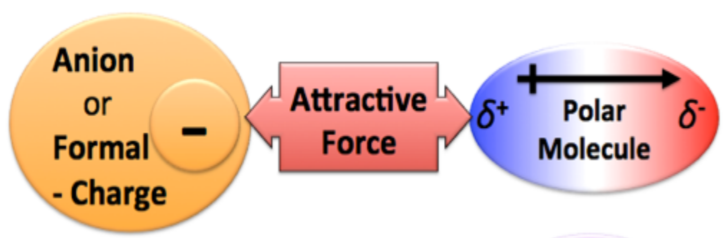
Understanding Check:

a. Can London dispersion forces occur between two CBr_4 molecules?
b. Can London dispersion forces occur between two H_2O molecules?
c. Which of the following substances is predicted to have the higher boiling point:

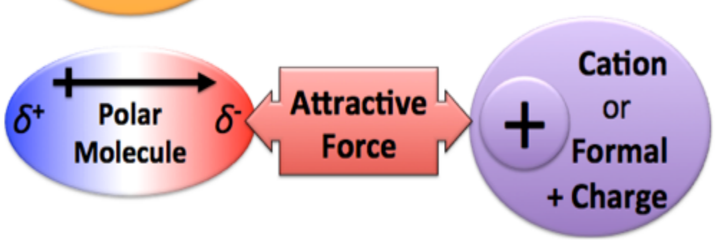
decane ($CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$)
or
nonane ($CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$)

4) Ion-Dipole Interactions

An **ion-dipole interaction**, as the name implies, is the electrostatic attractive interaction ***between*** an ***<u>ion</u>*** (or formal charge) ***and*** the ___________ of a polar molecule.

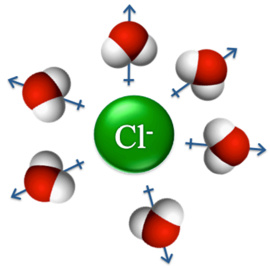


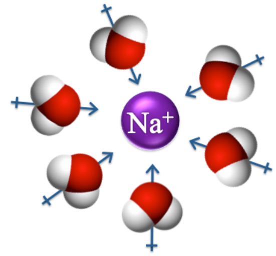
The attraction could be ***between*** an **anion** (or negative formal charge) ***and*** the partially positive end (**δ^+**) of a dipole, or vice versa, ***between*** a **cation** (or positive formal charge) ***and*** the partially negative end (**δ^-**) of a dipole.



The attraction could be ***between*** an **anion** (or negative formal charge) ***and*** the partially positive end (**δ^+**) of a dipole, or vice versa, ***between*** a **cation** (or positive formal charge) ***and*** the partially negative end (**δ^-**) of a dipole.

Examples: Sodium cations or chloride anions interacting with polar water molecules.





Understanding Check:

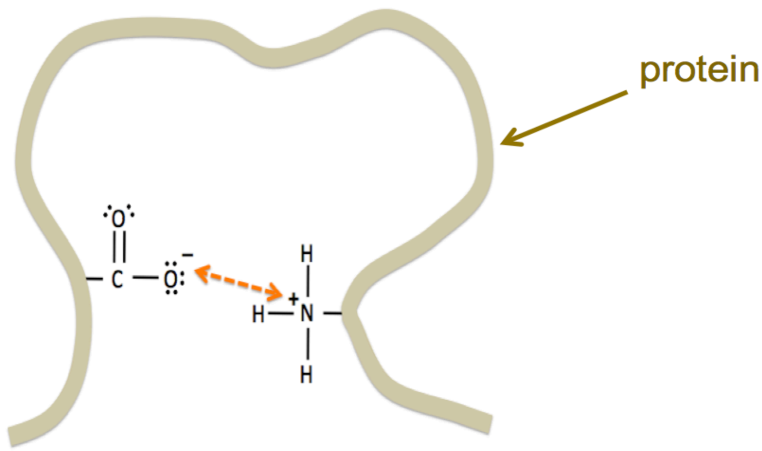
a. Can an ion-dipole interaction occur between ammonium (NH_4^+) and H_2O?

b. Can an ion-dipole interaction occur between bromide (Br^-) and I_2?

5) Salt Bridge Interactions

A **salt bridge** is the electrostatic attractive interaction ***between*** a negative formal charge ***and*** a positive formal charge in protein.

Salt bridges are one of the noncovalent interactions that are responsible for the way a protein ***folds up onto itself*** to give it the shape that is necessary for it to perform its biological function.



Summary of Noncovalent Interactions

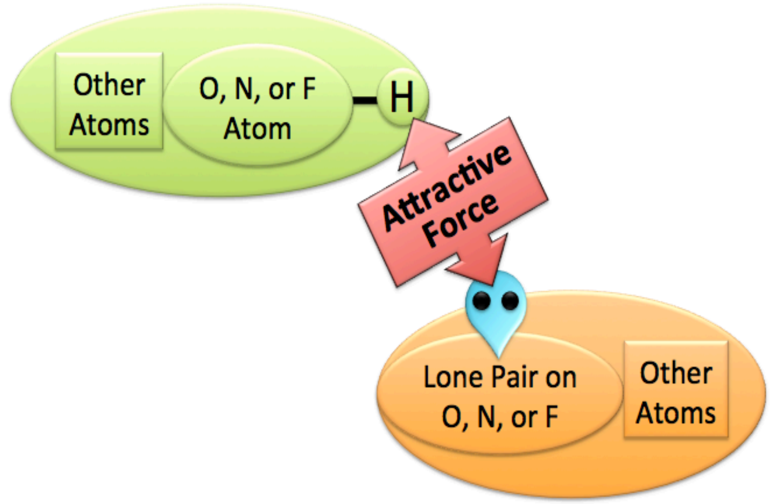
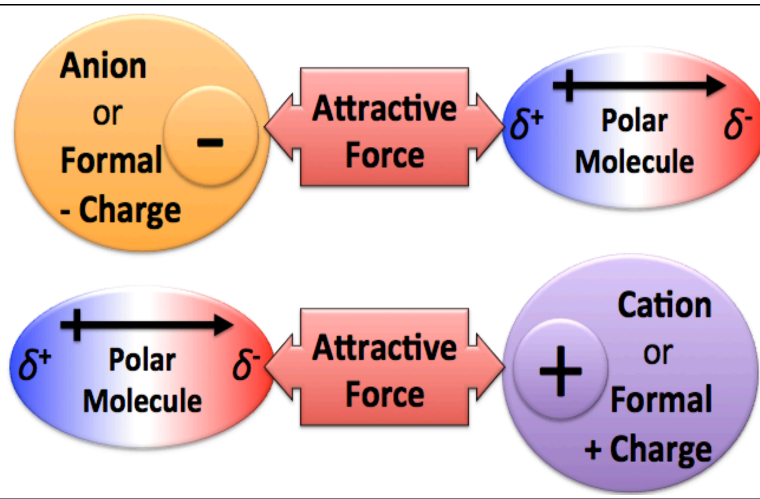
In this video, you learned about ***five*** types of **noncovalent interactions**.

When these interactions occur between molecules (as opposed to monatomic or polyatomic ions), they are sometimes referred to as **intermolecular forces**.

Noncovalent interactions cause molecules to be *attracted to each other* and can result in molecular compounds existing as liquids or solids.

In the case of large molecules or polyatomic ions that can *fold back upon themselves*, noncovalent interactions can occur between two regions ***within an individual particle***.

Summary of Noncovalent Interactions

Noncovalent Interaction	Interaction Between:
Hydrogen Bonding Electrostatic attractive force between the partially positive charged hydrogen end of an O-H, N-H, or F-H bond and the negative charge of a lone pair on an O, F, or N.	
Dipole-Dipole Electrostatic attractive force between *two* ***polar*** *molecules.*	
London Dispersion Forces Electrostatic attractive force between ***any*** *two molecules.*	
Ion-Dipole Electrostatic attractive force between a *dipole* and *an ion or formal charge.* The partially positive charge of a dipole is attracted to a negatively charged ion *or* a negative formal charge. The partially negative charge of a dipole is attracted to a positively charged ion *or* positive formal charge.	
Salt Bridge Electrostatic attractive force between a negative formal charge and a positive formal charge in protein.