Chemistry Unit 2

Primary reference: *Chemistry: Matter and Change* [Glencoe, 2017]

Topic	Essential Knowledge	Study and Practice
Scientific Investigation 1.2 **SOL 1b, 1g**	Understand and use **Material Safety Data Sheet** (MSDS) warnings including: handling chemicals, lethal dose (LD), disposal and chemical spill clean-up. The **percent by mass of an element in a compound** can be determined: % by mass of element = $\frac{\text{total mass of element in compound}}{\text{molar mass of the compound}}$ X 100	**Ch 10:** Read pp. 341-342 on percent composition
Atomic Structure and Periodic Relationships **2.2** **SOL 2a,2c,2d, 2e,2i**	**Democritus**:Greek philosopher who suggested the idea of atoms @ 400 BC. **John Dalton** atomic theory of 4 postulates was based on experimentation—early 1800s **J.J. Thomson** and **Millikan** discovered the electron and it's charge respectively. **Ernest Rutherford's** gold foil experiment showed the atom was mostly empty space with a small, dense, positively charge nucleus. Atoms are made of **protons, neutrons** in the nucleus. A cloud of electrons surrounds the nucleus. An atom's **atomic number** = the number of **protons**. All atoms of the same element have the same number of **protons**. A **proton** has a **positive** charge and a relative mass of one. The number of **electrons** equals the number of protons in a neutral atom. An **electron** has a **negative** charge and a relative mass of zero. A **neutron** has **no charge** and a relative mass of one. Isotopes are atoms of the same element with a different number of neutrons (Example C-12 and C-13). **mass number = #protons + # neutrons** The **atomic masses** on the periodic tables are a weighted average of the isotope masses. **Dmitri Mendeleev** created a Periodic Table based on the elements' masses and physical and chemical properties. **Moseley** reordered the table slightly based on atomic number. Rows are called **periods** and columns are called **groups** or **families**. Named families are **alkali metals, alkaline earth metals, halogens**, and **noble gases**.	**Ch 4:** Read pp. 102-105 on early atomic models and atomic theory. Read pp. 115-119 about atomic numbers and mass numbers **Ch 6:** Read pp. 174-181 about the periodic table.
Nomenclature, Formulas, and Reactions **3.2** **SOL 3a, 3b, 3c**	**Subscripts** in a **chemical formula** represent the relative number of each type of atom. The subscript follows the element symbol. Example: a water molecule, $\mathbf{H_2O}$, has 2 hydrogen atoms and one oxygen atom. **Parentheses** are used when a **subscript** affects a group of atoms. Example: $\mathbf{Mg(NO_3)_2}$ has a ratio of one magnesium atom, 2 nitrogen atoms and 6 oxygen atoms in the compound. **Molecules** form from non-metals and **ionic compounds** form from a metal cation and a non-metal anion. Metals lose electrons to become **cations**. Non-metals gain electrons to form **anions**. For ionic compounds, the charges of the anions and cations must add to zero. In binary ionic compounds, we name the metal first followed by the anion ending with –ide. Roman numerals are used to show the charge/oxidation state of metals other than alkali or alkaline earth metals. In binary molecular compounds, we use prefixes in front of the element names and end with –ide. A **chemical equation** shows the formulas of all the **reactants** on the left hand side of the arrow, and the formulas for all the **products** on the right hand side. Chemical reactions follow the **Law of Conservation of Mass**—matter is neither created nor destroyed during a chemical reaction. We **balance** chemical equations using **coefficients in front of each substance** in the equation so that each side has the same number of atoms of each element.	**Ch 3:** Read p. 85 **Ch 7:** Read pp. 206-209 on ions. *Electron configurations will be learned later.* Read pp. 210-216 on ionic compounds. Read pp. 221-224 on polyatomic ions and formulas pp 149-151 and 158-159 **Ch 9:** Read pp. 282-288 on chemical equations
Molar Relationships 4.2 **SOL 4a, 4b, 4d**	**Molar mass** is the sum of all the atomic masses in a compound. The **mole** can be used to convert between mass, particles and gas volume using unit cancelation. 1 mole = 6.02 x 10^{23} things = molar mass = 22.4 L(gas at 0°C & 1atm only) Ionic compounds **dissociate** in water to form **electrolyte** solutions (conduct electricity) whereas molecular compounds do not. An example of an ionic compound as it dissociates in water (into ions) is seen here: $MgBr_2(s) + H_2O(l) \rightarrow Mg^{2+}(aq) + 2Br^-(aq)$	**Ch 10:** Read pp. 325-340 on molar conversions; Read pp. 341-343 on percent composition (by mass). **Ch 7:** Read pp. 214-216 on electrolytes
Phases of Matter and Kinetic Molecular Theory **5.2** **SOL 5a, 5d**	**Kinetic Molecular Theory** describes the behavior of gases based on a model of an ideal gas. Ideal gases do not exist but help us understand how real gases behave. **Real gases** exist, have intermolecular forces, particle volume and can change states, whereas ideal gases do not. **Avogadro's hypothesis**: Equal volumes of gas at the same pressure and temperature will contain the same number of gas particles. 1 mole gas = 22.4 Liters at 0°C and 1 atm.	**Ch 12 and Ch 13:** Read pp. 400-406 for an introduction to gases; Read p. 452. *Gases will be revisited in greater detail;* Read pp. 457-459 on real and ideal gases.

Unit 2 Objectives

Chemistry: Matter and Change (Glencoe, 2017)

I. Basic Atomic Structure
   A. Early Atomic Models through Rutherford
   B. Atomic number, mass number, atomic mass and isotopes
II. Introduction to the Periodic Table
   A. Parts of the periodic table
III. Chemical Names and Formulas
   A. Differentiating between molecular and ionic compounds
   B. Ionic charges of Elements
   C. Names ↔ Formulas Binary Ionic Compounds
   D. Names ↔ Formulas Binary Molecular Compounds
   E. Diatomic Elements (Review)
IV. Mole Calculations
   A. Molar Mass
      1. Review of counting atoms in formulas
      2. Calculating molar mass
      3. Converting between moles and molar mass
   B. Molar Volume of Gases at STP
      1. Avogadro's hypothesis
      2. Converting between moles and molar volume at STP (1 mole gas = 22.4 L)
   C. More Molar Conversions
      1. Conversions: mass ↔ volume, mass ↔ count, volume ↔ count
V. Chemical Reactions
   A. Understanding chemical reaction symbols
   B. Balancing Chemical Reactions

Objectives (SOL)

1. Identify the contributions of Democritus, Dalton, Thomson, Rutherford, and Millikan, to the development of the modern atomic model. (2i)
2. Describe the structure of an atom, including the location of protons, electrons and neutrons.(2c).
3. Define the charges and relative masses of electrons, protons and neutrons.
4. Determine the number of protons, neutrons and electrons in elements and isotopes. (2a)
5. Explain how isotopes differ, yet are still the same element.(2a)
6. Calculate the atomic mass for an element given the weighted averages of the isotopes.(2b)
7. Identify the contributions of Mendeleev and Mosely to the modern periodic table.(2i)
8. Identify the following areas on the periodic table: alkali metals, alkaline earth metals, halogens, noble or inert gases, representative elements, transition metals, non-metals, metals, and metalloids.(2d)
9. Distinguish between ionic and molecular compounds.(2g, 3a)
10. Count the number of atoms present in compound formulas(3c)
11. Explain how anions and cations are formed.(2g)
12. Predict monatomic ion charges using the periodic table (2g)
13. Use the roman numeral Stock System to identify and name transition metal ions.(3a)
14. Predict the ionic compound formed from any two monatomic ions.(3c)
15. Write the formulas for binary ionic and molecular compounds given their names and *visa versa*.(3a&3c)
16. Name the seven diatomic elements.(3a)
17. Explain Avogadro's Hypothesis.(4a)
18. Memorize molar volume = 22.4 Liter at 1 atmosphere and 0°C(4a)
19. Calculate the molar mass of a substance given the formula.(4a)
20. Calculate conversions between moles, molar masses, molar volumes, and particle counts.(4a)
21. Master reading and writing chemical equations using chemical formulas and symbols correctly. (3b)
22. Explain the Law of Conservation of Mass
23. Balance equations (3b)
24. Explain a catalyst's role in a chemical reaction. (3f)

Chapter 5 Atomic Structure Skeleton Notes

What is an <u>atomic number</u> of an element and where do we find it?	
What is a <u>mass number</u>?	
What is an <u>isotope</u>?	
How do you read isotope symbols? $^{6}_{3}Li$ or ^{6}Li or Li-6 *vs* $^{7}_{3}Li$ or ^{7}Li or Li-7	top # = ______________________ bottom # = ____________________________
How many protons, neutrons and electrons are in $^{35}_{17}Cl$?	protons ______, neutrons______, electrons ________
How many protons, neutrons and electrons are in $^{37}_{17}Cl$?	p^+ ________, n^0________, e^- ___________
How many protons, neutrons and electrons are in Calcium-42?	p^+ ___________, n^0___________, e^- _________
What is the atomic mass of an element and where do we find it?	

Calculating Atomic mass

Elements contain a mix of isotopes. If we are given the percent composition of each isotope, we can calculate the atomic mass using weighted averages.

Analogy: Weighted Grades

Type	% Weight	x Score	= Contribution
Tests	50%	75%	
Quizzes	25%	92%	
Homework	25%	95%	

We use mass number x % abundance (composition) to calculate the approximate atomic mass
Example: Find the atomic mass of chlorine using the data below. (Ans = 35.4846 amu)

Isotope	% Abundance	x mass number	= Contribution
Cl-35	75.77	35	
Cl-37	24.23	37	

We can calculate a more accurate value by using % abundance and isotope mass in atomic mass units, amu, to calculate the value. (Ans = 35.4528)

Isotope	% Abundance	x amu	= Contribution
Cl-35	75.77	34.969	
Cl-37	24.23	36.966	

Practice: Naturally occurring oxygen contains 99.757 % Oxygen-16, 0.038% Oxygen-17 and 0.205% Oxygen-18. Calculate the approximate atomic mass.(Ans = 16.00448 amu)

Practice: Use the atomic mass unit data in the following table to calculate oxygen's atomic mass more accurately. (Ans = 15.999 amu)

Isotope	% abundance	amu
O-16	99.757	15.995
O-17	0.038	16.999
O-18	0.205	17.999

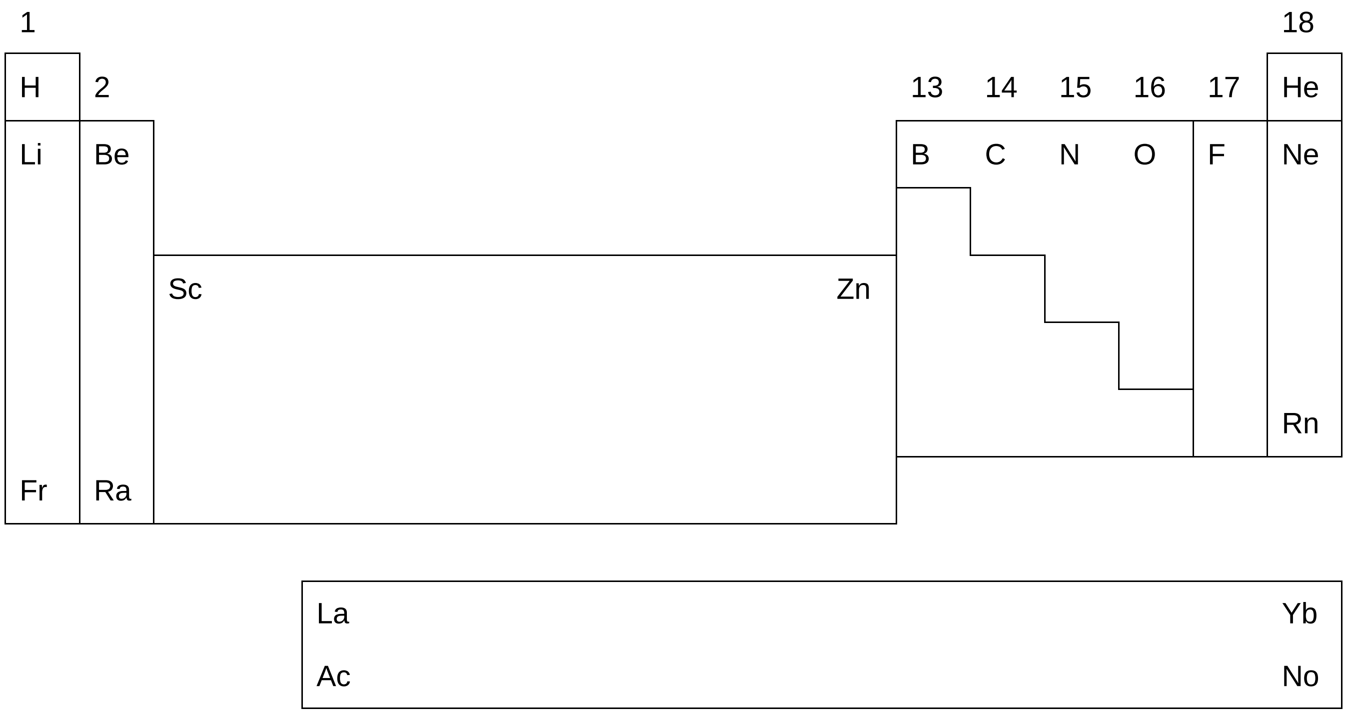
Animation for mass spec and isotopes at:
http://wps.prenhall.com/wps/media/objects/4974/5093961/emedia/ch02/MassSpectrometer/c2s4item20/MassSpectrometer.html

Chapter 6: Chemical Names and Formulas

Areas of the Periodic Table (Chapter 5)

Columns are called ________________ or ____________________

Rows are called _________________________



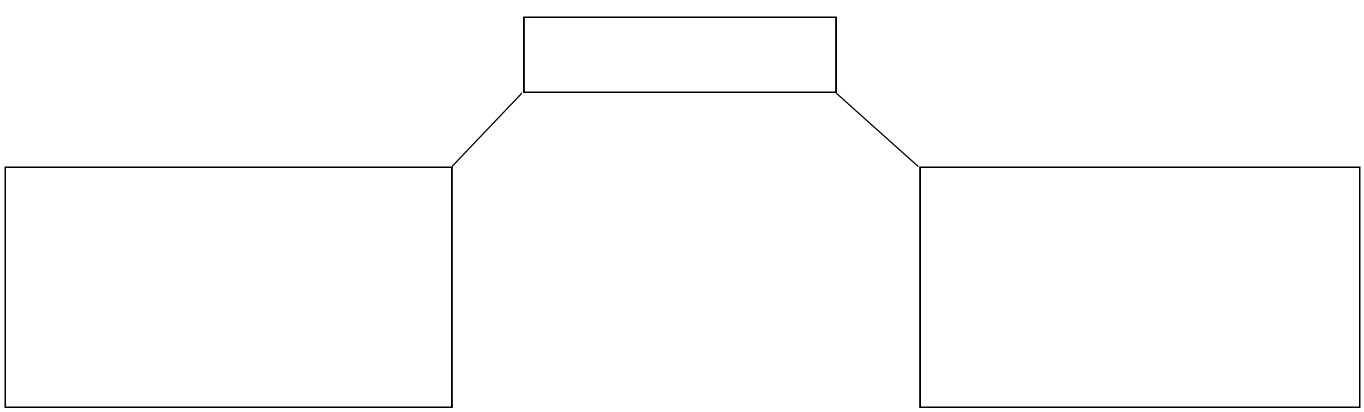
Nonmetals are located in the _________________________

Metals are located in the ____

A. Types of Compounds

1) Compounds are ____ elements

2) Elements may combine to form __________________ or ________________compounds



3) IONIC COMPOUNDS form when a ___________________________ combines with a

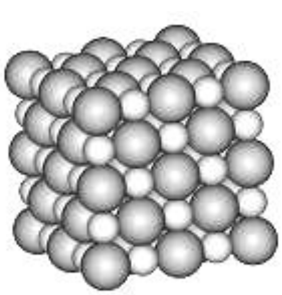
____

a) cations are metal atoms that have ______________ electrons, so they acquire a

_________________ charge

b) anions are nonmetal atoms that have __________________ electrons so they acquire

a ______________________ charge



Formula Unit:____

____

Properties of Ionic Compounds

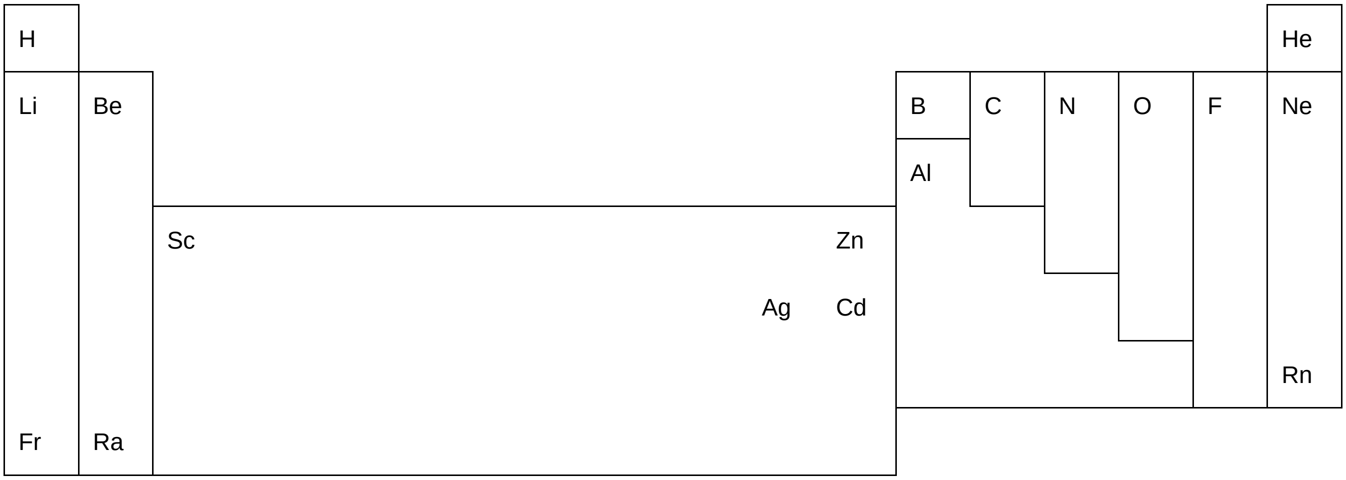
Melting points are _______________________________________

Physical State at room temperature_________________________

Ionic Compounds dissociate into ions in water when they dissolve.

B. Common Oxidation States (Charges) of Elements in Ionic Compounds

1. For alkali metals, alkaline earth metals, non-metals and noble gases.



2. For all other elements

- Transition metals may have more than one common ____________________________
- The oxidation state is indicated by a ____
- Example: iron(II) = ___________ iron(III) = ________________
- Exceptions: silver _______, Zinc ________. Cadmium _______, Aluminum _______

Review: formulas tell us two things:

1) ____

2) ____

Examples:

$NaCl$ Al_2O_3 $Al_2(SO_4)_3$ $CH_3(CH_2)_{12}CH_3$

C. Binary Ionic Compounds

1) Binary compound:____

2) Ionic compounds are ____________________________. The positive charges from the cation must be ______________________ by the negative charges from the anion.

3) The cation (metal) is always written first, and the anion ends in _____________.

Examples: sodium chloride is NaCl

Iron(III)oxide is Fe_2O_3

4) Steps for writing Binary Ionic Formulas from names:

a) ____

____

b) ____

____

Examples:

Calcium bromide

copper(II)chloride

Magnesium nitride

lead(IV)sulfide

Zinc phosphide

Lithium oxide

5) Steps for naming binary ionic compounds from formulas with transition metals

a) ____

b) ____

c) ____

d) ____

____

K_3N SnF_2 Fe_2O_3

TiO_2 CaO

D. Binary Molecular Compounds

1) Composed of two ______________________

2) Because they are molecules, no ______________________ are involved.

3) Prefixes are used to show the number of each type of atom present:

1 = ______________	6 = ______________
2 = ______________	7 = ______________
3 = ______________	8 = ______________
4 = ______________	9 = ______________
5 = ______________	10 = ______________

4) Exceptions: Do not use mono ______________________________

Drop the prefixes ending –a or –o with oxygen compounds.

Examples:

Phosphorus trichloride ______________________________

Nitrogen dioxide ___________________________________

Dinitrogen pentoxide ________________________________

Carbon monoxide __________________________________

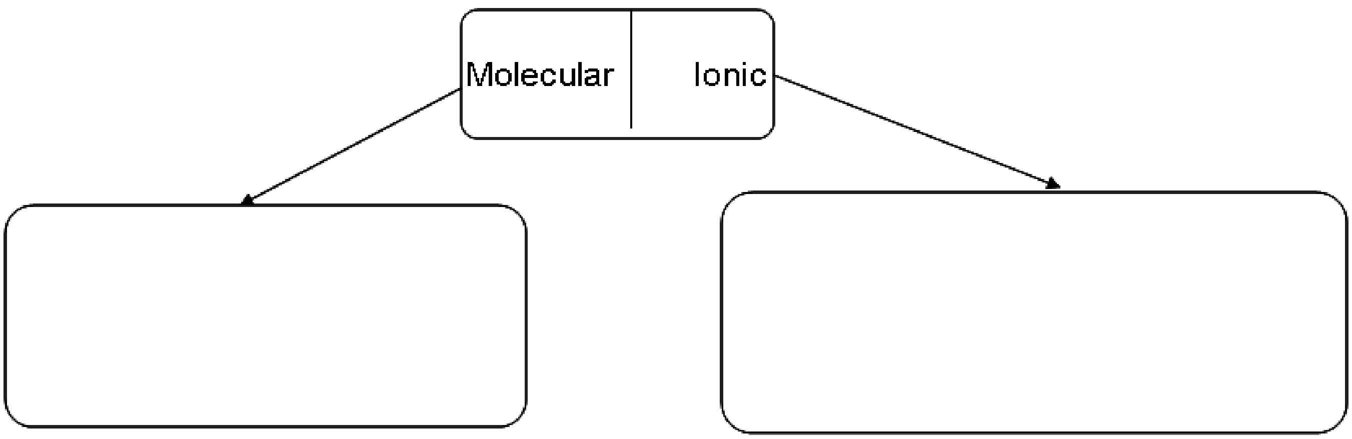
CF_4 _____________________________________

P_2O_5 ____________________________________

SiO_2 ____________________________________

$AsCl_3$ ____________________________________

Mixed Ionic and Molecular Naming



Circle the molecular compounds, then name all compounds correctly.

CaF_2 $\qquad$ CaO

NF_3 $\qquad$ P_2O_5

SiO_2 $\qquad$ Fe_2N_3

E. DIATOMIC ELEMENTS: Professor _________________________

The Mole & Avogadro's Number

1 mole of ANYTHING (cars, people, atoms, molecules, books, protons…) is equal to ________________________. You definitely need to memorize this.

1 mole is equal to ↑, just like *1 dozen is equal to 12 things*.

Moles are usually referring to atoms and compounds (**1 mole of Na atoms** = 6**.022 × 10^{23} Na atoms!**) because atoms and compounds are very, very, very small.

_______________________ is Avogadro's number. It's equal to 1 mole. (you only need to know it to "6.02"…). GET OUT YOUR PERIODIC TABLE.

Video Guide: "How Big is a Mole?" (TED-Ed)

1) Who was the first guy to propose numerical "counting" of particles like atoms and molecules?

2) If you have 6.02 × 10^{23} (that's about 602,000,000,000,000,000,000,000) molecules of water (H_2O)…
   a. …how much will it weigh in grams? ___________g
   b. …and since the density of H_2O is 1 g/mL, 18.01 g of H_2O should also have a volume of ___________mL.

KNOW THIS:

1 mole of He atoms = 6.02 × 10^{23} He atoms.
1 mole of P atoms = 6.02 × 10^{23} P atoms
1 mole of Cu atoms = 6.02 × 10^{23} Cu atoms.
1 mole of Na atoms = 6.02 × 10^{23} atoms of Na.
1 mole of H_2 molecules = 6.02 × 10^{23} molecules of H_2.
1 mole of CO_2 molecules = 6.02 × 10^{23} molecules of CO_2. EASY!

Check out the next pattern. Have your PT ready:

1 mole of He atoms = 6.02 × 10^{23} He atoms = **4.003 g of He**
1 mole of P atoms = 6.02 × 10^{23} P atoms = **30.97 g of P**
1 mole of Cu atoms = 6.02 × 10^{23} Cu atoms = _____________ **g of Cu**
1 mole of Na atoms = 6.02 × 10^{23} atoms of Na = _____________ **g of Na**
1 mole of H_2 molecules = 6.02 × 10^{23} molecules of H_2 = **2.02 g of H_2**

The first tricky thing: How many H atoms are in each molecule of H_2? ______

So… you can easily find the **molar mass** of H on the PT. Why did we have to multiply that PT number (for a *single* H) by **2** in order to get the **molar mass** of H_2?

1 mole of CO_2 molecules = 6.02 × 10^{23} molecules of CO_2 = _______________ g of CO_2.
(1 carbon is ___________ **g/mol**. 2 oxygens = 2×_____________ **g/mol**)

KNOW THIS before you practice: When you look on the PT and see that the **molar mass** for magnesium (**Mg**) is ____________, you've GOT to have units!

The unit for a regular pack of eggs is 12 **eggs/dozen**.
The unit for an average pack of paper is 500 **sheets/ream**.
The unit for a molar mass is **grams/mole**. (seen usually as **g/mol** for short.)

Worked Example for Calculating the Molar Mass of H_2O:

Worked Example for Calculating the Molar Mass of **nitrogen trifluoride**:

Worked Example for Calculating the Molar Mass of **ammonium chloride**:

Practice. Find the molar mass of each compound, using your PT. Must use appropriate units!

1) How much molar mass does 1 mole of zirconium have?
2) Determine the molar mass of 1 mole of sodium chloride.
3) Find the molar mass of 1 mole of CF_4.
4) What is the molar mass of 1 mole of barium hydroxide?

*****Always start with what you know*****

5) Worked Example: I have 1.95×10^{24} atoms of sulfur. (Sig figs matter!)
   a. How many moles of sulfur do I have?
   b. How many grams does my sample weigh?

6) Worked Example: Mark has 88.0 grams of solid KF.
   a. How many moles of KF does Mark have?
   b. How many formula units (particle-pairs/set) of KF does Mark have?

<u>Conversion Mapping:</u>

You can go from ______________ to _____________ (and vice-versa) by using the molar _________, found on the period table. ***You'll have to do some adding and multiplying for compounds.***

Review: What are the units for molar mass?

Review: How would you calculate the molar mass for K_2O, knowing there are ___ K atoms and ___ O atom?

You can go from ______________ to _________________ (and vice-versa) by using _________________'s Number.

Review: What is Avogadro's #?

Regardless of what you do in chemistry, **you MUST go <u>through</u> the mole**:

Mixed Practice:

7) How many molecules of water are in 60.0 grams of it?

8) How many grams does 4.77×10^{19} **formula units** of NaBr weight?

9) Which element has a molar mass of 196.97 g/mol? ________________________

10) Which element has a molar mass of about 83.8 g/mol? ________________________

11) Which **diatomic** element has a molar mass of 37.997 g/mol? ____________________

12) 6.02×10^{23} atoms of vanadium will have a mass of ________________ grams.

13) 7.5593×10^{38} atoms of vanadium will have a mass of _______________ grams.

14) 7.5593×10^{38} atoms of vanadium is equal to ______________ moles of vanadium.

15) 3.7 moles of H_2O will have a mass of ________________ grams.

16) 32,600 milligrams of carbon will contain ____________________ C atoms.

17) 10.0 moles of phosphorus pentafluoride will have a mass of _________________ grams, and will contain ______________________ molecules of the compound.

18) Students calculated the molar mass of strontium iodide to be 307.13 g/mol. Calculate their % error.

GASES and The Mole

Gases take up space. A lot of space!

Recall that the amount of _________________ in something give us its ***mass***, and the amount of ***space*** that something takes up is called __________________.

Volume has a few different units.

We know that 1 cm^3 = _____________ = 1 cc.

… and 1000 mL = 1 ______________

The common unit for VOLUME in chemistry is the ________________.

VERY IMPORTANT:
1 mole of a gas is gonna take up __________________.
In fact, 1 mole of any gas will take up _______________ L of space.

KNOW THIS: 1 mole (g) = 22.4 L

Review: According to Professor BrINClHOF ("Brinklehoff"), what are the 7 diatomic elements? Write them out in X_2 format, and circle the ones that are gases at standard (normal) temperature and pressure.

Practice with moles and gases: Use the right units! ALWAYS GO THROUGH THE MOLE!

19) How much volume does 8.40 moles of hydrogen (H_2) gas have?

20) How much space (in L) does 113 moles of O_2 take up?

21) 5.4×10^{16} atoms of helium gas will take up _______________ liters of space.

22) How much mass does 59.4 L of chlorine gas have?

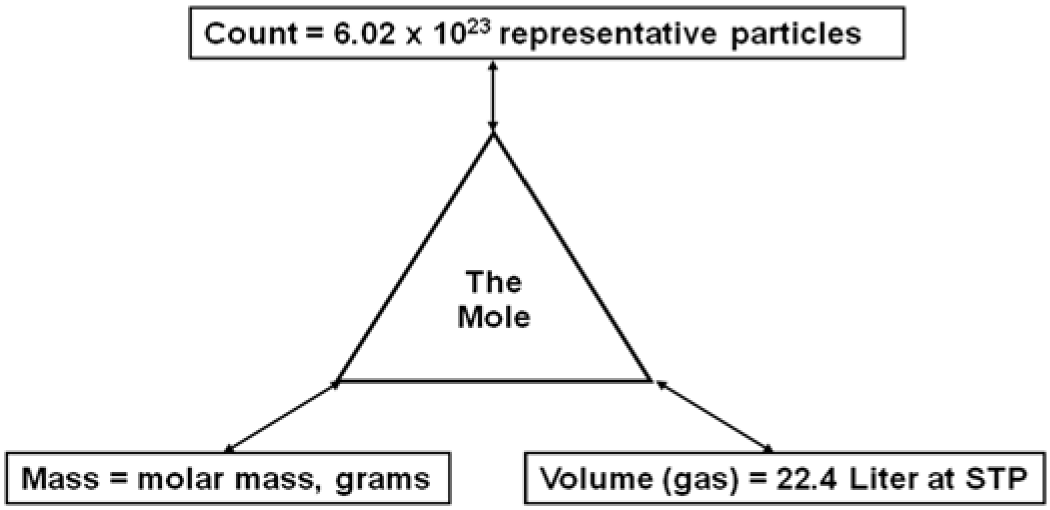
23) 44.8 L of nitrogen gas is _____________ moles of nitrogen.

24) 44.8 L of nitrogen gas is _____________ grams of nitrogen.

25) 44.8 L of nitrogen gas is _____________ molecules of nitrogen.

26) Determine the mass of carbon dioxide gas if the volume of the gas is 96.7 L.

Chapter Seven: The Mole is the Chemist and Physicist's Dozen



The mole = the number of atoms in exaxtly twelve grams of Carbon-12.

1958: 1 mole = 6.02×10^{23}
2006: 1 mole = $6.02214078 \times 10^{23}$

Avogadro's number = ________________________

Some mole facts—see if you can find the pattern

- 6.02×10^{23} atoms in twelve grams of C-12.
- 6.02×10^{23} atoms in 12.011 g of carbon (naturally occurring
- 6.02×10^{23} atoms in 1.008 g of hydrogen
- 6.02×10^{23} atoms in _________ g of oxygen

Converting between moles and counts of representative particles

1 mole = 6.02×10^{23} representative particles or

$$\frac{1\text{ mol}}{6.02 \times 10^{23}\text{ rep. part}} = \frac{6.02 \times 10^{23}\text{ rep. part.}}{1\text{ mol}}$$

Representative particles:

Elements____

Molecules ____

Ionic compounds____________________________________

$\boxed{\text{1 mole} = 6.02 \times 10^{23} \text{ representative particles}}$

1. How many formula units of MgS are there in 0.482 mol of MgS?

2. How many moles are in 1.204×10^{25} molecules of nitrogen dioxide?

3. How many sodium atoms are there in 3.2 moles of sodium?(Ans = 1.9×10^{24} Na atoms)

4. How many moles are there in 6.32×10^{24} formula units of Iron(III) sulfide?(Ans = 10.5 mol)

Review of Counting Atoms in Formulas

One mole of NaCl = ______________________________

One mole of Na_2S = __________________________

One mole of $Al_2(SO_4)_3$ = ____

You do:

Li_3PO_4 = ____

$Pb(CO_3)_2$ = ____

$Fe_2(Cr_2O_7)_3$ = ____

Calculating Molar Mass (formula units, atoms, molecules)
Definitions:

- Molar Mass = the mass of $6.02x10^{23}$ representative particles of an element, molecule, or ionic compound. Molar Mass may also be called Formula Mass.

6.02×10^{23} representative particles=1 Mole = molar mass, gram

Examples:

1. Find the molar mass of Sodium.
2. Find the molar mass of NaCl.
3. Find the molar mass of $CaCl_2$
4. Find the molar mass of $Cu(NO_3)_2$
5. Find the molar mass of $Mg(OH)_2$
6. Find the molar mass of $Ca_3(PO_4)_2$

Converting between Moles and Molar Mass

1 mole = 6.02 x 10^{23} rep. particles = molar mass, g

$$\frac{1 \text{ mol}}{\text{molar mass, g}} = \frac{\text{molar mass, g}}{1 \text{ mol}}$$

1. What is the mass of 2.3 moles of $MgBr_2$?(Ans = 420 g)

2. How many moles of potassium iodide are in 29.3 g of KI? (Ans = 0.177 mol)

3. How many grams of SO_3 are present in 2.3 moles of SO_3?(Ans = 180 g)

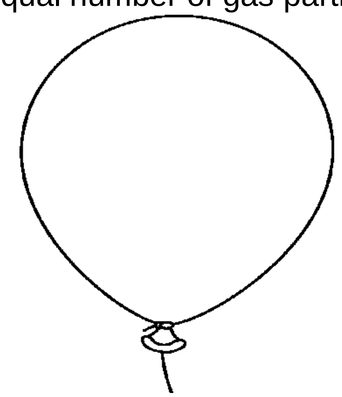
4. How many moles of CaF_2 are equivalent to 450 grams of CaF_2?(Ans = 5.8 mol)

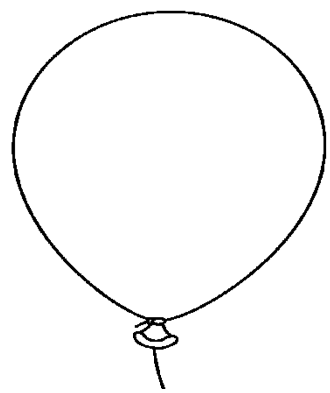
5. How many grams of titanium(IV) sulfide, TiS_2, are present in 0.056 moles of titanium(IV) sulfide? (Ans = 6.3 g)

6. How many moles of ammonium sulfate $(NH_4)_2SO_4$ are present in 52.3 grams of ammonium sulfate?(Ans = 0.396 mol)

Molar Volumes

Avogadro's Hypothesis: Equal volumes of gas at the same temperature and pressure contain an equal number of gas particles (atoms or molecules).





Molar Volume: The volume of one mole of gas at standard temperature and pressure (STP)

STP = ____

$\boxed{\text{1 mole gas = 22.4 L at STP for any gas}}$

Assumptions about ideal gases

- The gas particles have no volume (points in space)
- The gas particles have no intermolecular attractions
- The gas particles collide elastically like billiard balls.
- Ideal gases never condense no matter how cold it is.

Real Gases Condense!!!!

Converting between Moles and Molar Gas Volume

1 mole = 6.02×10^{23} rep. particles = molar mass, g = molar gas volume (22.4L at STP)

$$\frac{1\ \text{mol}}{22.4\ \text{L}} = \frac{22.4\ \text{L}}{1\ \text{mol}}$$

1. A neon light contains (neon is a noble gas) 0.51 liters of neon gas at STP. How many moles does the light contain?(Ans = 2.3×10^{-2} mol)

2. A helium balloon contains 0.325 moles of gas at STP. What is the balloon's volume in liters? (Ans = 7.28 L He)

3. An underground cavern contains 5.5×10^5 liters of natural gas, CH_4, at STP. How many moles of gas are in the cavern? (Ans = 2.5×10^4 mol)

4. A blimp contains 35,000 liters of hydrogen gas (flammable!) at STP. How many moles of hydrogen does it contain? (Ans = 1.6×10^3 mol)

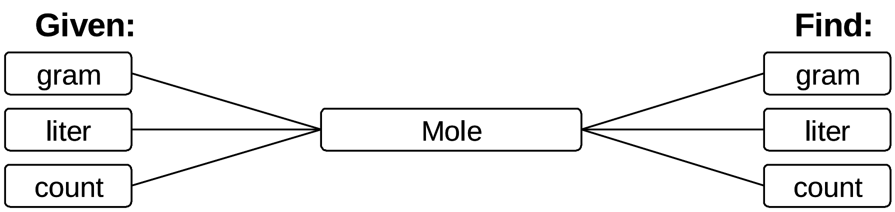
More Molar Conversions—You can go anywhere!

1 mole = 6.02 x 10^{23} rep. part = gram formula mass = 22.4 Liters at STP

Chapter Seven Objectives

- Memorize Avogadro's number
- Memorize STP volume of 1 mole of gas (0°C, 1 atm)
- Convert between grams, moles, representative particles and liters using factor label method.
- Calculate molar masses

unit to mole	**mole to unit**
$\frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ rep. part.}}$	$\frac{6.02 \times 10^{23} \text{ rep. part.}}{1 \text{ mole}}$
$\frac{1 \text{ mole}}{\text{g. molar mass}}$	$\frac{\text{g. molar mass}}{1 \text{ mole}}$
$\frac{1 \text{ mole}}{22.4 \text{ L. gas}}$	$\frac{22.4 \text{ L. gas}}{1 \text{ mole}}$



1. How many atoms of gold are contained in a 505 gram bar of Au?(Ans = 1.54 x 10^{24} Au atoms)

2. Find the number of moles of Cl_2 gas in a 1.46 x 10^4 liter tank at STP.(Ans = 62 mol)

3. A balloon contains 1.2 grams of Helium. What is the balloon's volume at STP?(Ans = 7.3 L)

4. How many molecules of fluorine gas are in an 0.0030 liter ampule at STP?(Ans = 8.2 x 10^{19} molec.)

Chapter Eight Skeleton Notes Part 1

I. Chemical Equations

Symbols	Use
$\rightarrow$	
$\leftrightarrows$	
(s), (l), (g)	
(aq)	
$\xrightarrow{catalyst}$	
$\xrightarrow{heat}$	

B: Examples of chemical equations

- $Mg(s) + O_2(g) \rightarrow MgO$
  Reactants = Products =

- $OF_2(g) \rightarrow F_2(g) + O_2(g)$
  Reactants = Products =

Writing chemical equations from word equations

1. Sodium metal reacts with chlorine gas to form sodium chloride

2. Iron metal reacts with oxygen gas to form rust, iron(III)oxide.

3. Solid nitrogen triiodide decomposes to solid iodine and nitrogen gas.

Catalysts:____

____

Skeleton equations do not show the amounts of products and reactants.

C. Balancing chemical reactions using coefficients

Law of conservation of mass: ____

____

Balanced equations use coefficients in front of formulas to show the _______________

____

A chemical reaction is balanced if there are the same number of each kind of element on both sides of the chemical equation. If there are four oxygens on the reactant side, there will be ____ oxygens on the product side.

Count the atoms of each element of both sides.→ Indicate which equations are balanced.

$C(s) + O_2(g) \rightarrow CO_2(g)$

$Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

$3NH_3 + H_3PO_4 \rightarrow (NH_4)_3PO_4$ coefficients apply to the entire compound

$C_4H_{10} + 4O_2 \rightarrow 4CO_2 + 5H_2O$

$Ca(ClO_3)_2 \rightarrow CaCl_2(s) + 3O_2(g)$

balancing chemical equations

a. Write the skeleton chemical equation leaving blanks for the coefficients:
b. Count the number of each element in the reactant and product side
c. Balance the equation using whole number coefficients (NEVER SUBSCRIPTS)
d. Track your changes
   1. Balance the other compounds to the most complicated compound.
   2. Balance the binary compounds (H_2O, CO_2, NO_2)
   3. Balance diatomics and elements last
   4. If you end up with an odd number that won't balance (3 oxygens on one side, two on the other) double all the coefficients filled in so far.
   5. Double check when you're done.

$\_\_\_N_2O_4(g) \xrightarrow{Pt} \_\_\_N_2(g) + \_\_\_O_2(g)$

$\_\_\_K_2S + \_\_\_FeCl_3 \rightarrow \_\_\_Fe_2S_3 + \_\_\_KCl$

For combustion reactions, use the CHO rule (C first, H second, O last)

$\_\_\_CH_4 + \_\_\_O_2 \rightarrow \_\_\_CO_2 + \_\_\_H_2O$

$\_\_\_C_2H_6 + \_\_\_O_2 \rightarrow \_\_\_CO_2 + \_\_\_H_2O$

$\_\_\_C_2H_6O + \_\_\_O_2 \rightarrow \_\_\_CO_2 + \_\_\_H_2O$

Chlorine gas was used in chemical warfare during WWI. The Germans used Chlorine gas on the Allied Forces in Ypres, France in 1915. Chlorine reacts with the moisture in lungs to produce hydrochloric acid, HCl.

You try balancing the reaction for chlorine in your lungs:

___$Cl_2(g)$ + ___$H_2O(\ell)$ → ___$HCl(\ell)$ + ___$O_2(g)$

Now try balancing the reaction for phosgene (Cl_2CO) in your lungs. This is another poisonous gas used in warfare.

___$Cl_2CO(g)$ + ___$H_2O(\ell)$ → ___$HCl(\ell)$ + ___$CO(g)$ + ___$O_2(g)$