**Review for Mid-Year Exam**

**Learning Standards:**

**HS-PS1-1.** Use the periodic table as a model to predict the relative properties of main group elements, including ionization energy and relative sizes of atoms and ions, based on the patterns of electrons in the outermost energy level of each element. Use the patterns of valence electron configurations, core charge, and Coulomb’s law to explain and predict general trends in ionization energies, relative sizes of atoms and ions, and reactivity of pure elements.

**HS-PS1-2.** Use the periodic table model to predict and design simple reactions that result in two main classes of binary compounds, ionic and molecular. Develop an explanation based on given observational data and the electronegativity model about the relative strengths of ionic or covalent bonds.

* Simple reactions include synthesis (combination), decomposition, single displacement, double displacement, and combustion.
* Predictions of reactants and products can be represented using Lewis dot structures, or chemical formulas.
* Observational data include that binary ionic substances (i.e., substances that have ionic bonds), when pure, are crystalline salts at room temperature (common examples include NaCl, KI, Fe2O3); and substances that are liquids and gases at room temperature are usually made of molecules that have covalent bonds (common examples include CO2, N2, CH4, H2O, C8H18)

**HS-PS1-7.** Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.

* Mathematical representations include balanced chemical equations that represent the laws of conservation of mass and calculations of percent yield.

**OUTLINE OF TOPICS WE HAVE COVERED**

* Basic atomic structure
* Bohr Model of atoms
* Isotopes
* Valence electrons
* Drawing Lewis dot structures
* Organization of the Periodic Table
* Classification of elements (metals, nonmetals, metalloids)
* Families of elements
* Ionization Energy
* Relative atomic and ionic size
* Electronegativity
* Reactivity
* Ion formation
* Lewis structures
* Formation of ionic compounds
* Ionic bonding properties and formulas
* Naming and writing formulas of ionic compounds
* Properties of Ionic Compounds
* Naming and writing formulas of covalent compounds
* Lewis structures
* Properties of Covalent Compounds
* Diatomic elements
* Chemical change vs physical change
* Types of Reactions
* Writing and Balancing Equations
* Molar Mass and Mole Conversions (moles, mass, volume, # particles)
* % Composition
* Empirical Formulas
* Molecular Formulas

The **periodic table** contains information for every element.

**At. # = 6 Atomic Mass = 12.01 amu**

**Mass # = 12**



*Define the terms atomic number, atomic weight, and mass number. Identify the atomic number, atomic weight, and mass number for carbon based on the information on your periodic table.*

**At. #** - the # of protons in the nucleus of an atom of an element. (Identifies the element) Also the # of e-s in a neutral atom. **Atomic weight (mass)** – the average mass of all the naturally occurring isotopes of an element (found on the periodic table) **Mass number** – the sum of the protons and neutrons in the nucleus of an isotope of an element.

***Atomic structure***for any element can be determined from the information on the periodic table.

*Describe the three subatomic particles, their charges, sizes, and locations in an atom.*

*Determine the number of protons, neutrons, and electrons for a Boron atom.*

**Proton (p+)** +1 charge, 1 amu, in the nucleus **Neutron (n0)** no charge, 1 amu, in the nucleus **Electron (e-)** -1 charge, 1/1800 amu, surrounding the nucleus in energy levels. **Boron:** 5 p+, 6 n0, 5 e-

***Bohr Models*** represent the structure of the atom with protons and neutrons in the nucleus and electrons on “rings” outside of the nucleus. Remember the first ring holds 2, the second ring holds 8, and the third ring holds 18.

 

 *Draw the Bohr models for Na and Cl and Ar*

***Isotopes*** are atoms of the same element because they have the same number of protons but a different number of neutrons.

 *Determine the number of protons, neutrons and electrons in Hydrogen-1, hydrogen-2, and hydrogen-3*

*All have 1 p+ and 1 e- Hydrogen-1 has no neutrons, Hydrogen-2 has 1 neutron, and Hydrogen-3 has 2 neutrons*

The last digit of an element’s group number is equal to its ***number of valence electrons***.

 *Which contains the greatest number of valence electrons?*

 *Ca Ge Se Kr*

Draw one dot for each valence electron when drawing an element’s or ion’s ***Lewis electron dot diagram***.

 *Which dot model would contain the fewest dots as valence electrons?*

 *Ca Ge Se Kr*

There are various trends that can be found on the periodic table. Define **atomic radius, electronegativity, ionization energy,** and **electron affinity**. Identify their trends using arrows and the words “increasing” or “decreasing” on the periodic table below.



The three classes of elements on the periodic table are **metals, nonmetals,** and **metalloids.** Describe their general characteristics. Color all of the metals green, the metalloids red, and the nonmetals blue on the table above.

Atoms are most stable when they have 8 valence electrons (an ***octet***) and tend to form ions to obtain such a

 configuration of electrons.

 *Predict the charge of the ions formed by these elements*

 *Li 🡪* ***Li+*** *F 🡪* ***F-***  *Na 🡪* ***Na+*** *Cl 🡪* ***Cl-***

**Chemical formulas** are written so that the charges of **cations** and **anions** neutralize (cancel) one another.

 *calcium phosphate*: Ca2+ PO43- = **Ca3(PO4)2** (criss-cross)

When naming **binary ionic compounds**, write the name of the positive ion (cation) first,

followed by the name of the negative ion (anion) with the name ending in “-ide.”

*CaCl2* ***calcium chloride*** *MgS* ***magnesium sulfide***

When naming compounds containing **polyatomic ions**, keep the name of the

 polyatomic ion the same as it is written on your reference sheet.

 *NH4Cl* ***ammonium chloride*** *copper (I) nitrate* ***CuNO3***

**Roman numerals** are used to show the positive oxidation number of the cation if it has more than

 one positive oxidation number

 *FeO:* ***Fe2+ O2-*** *so* ***iron(II) oxide*** *Nickel (III) sulfate:* ***Nii3+ SO42- so Ni2(SO4)3***

***Ionic bonds*** form when one atom ***transfers*** an electron to another atom when

 forming a bond with it.

 *Which substance exhibits ionic bonding rather than covalent bonding?*

######  CO2 N2O4 SiO2 **CaBr2** C6H12O6

***Ionic bonds*** form when the electronegativity difference between two bonding atoms is *great enough to cause the transfer of electrons.*

 *Which of the following combinations would form an ionic bond?*

 *H and H* ***Na and N*** *H and N* ***Na and Br***



***Covalent bonds*** form when two atoms ***share*** a pair of electrons.

 *How many covalent bonds are found in a nitrogen (N2) molecule?*

 **A triple covalent bond (3 bonds between the 2 N atoms)**

Substances containing mostly covalent bonds are called ***molecular substances***.

They are attracted to each other by sharing electrons

*Which of the following is a molecular substance?*

*Lithium chloride* ***carbon monoxide*** *sodium nitrate aluminum oxide*

***Lewis*** ***Electron dot structures*** represent covalently bonded molecules formed through the sharing

 of valence electrons between atoms.

 *Draw the Lewis structures for these common molecules H2 H2O NH3 CH4*

   

When naming a **binary molecular compound,** the first element is named using the name of the element.

 The second element always end in “–ide.” Indicate the number of atoms using the prefix…

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

If the first element has only one atom, don’t use the mono.

*What is the name of the following molecular compounds?*

 *CO2 \_****carbon dioxide***  *N2O­­­­­­­­­­­­­­­\_****dinitrogen monoxide***

Complete and memorize this table

|  |  |
| --- | --- |
| Substance Type | Properties |
| **Ionic** |  |
| **Covalent (Molecular)** |  |

***Physical changes*** do not form new substances.

They merely change the appearance of the original material. (The melting of ice) H2O (s) 🡪 H2O (l)

***Chemical changes*** result in the formation of new substances or the product(s) of a ***chemical reaction***.

 *Which process is an example of a chemical change?*

*the melting of ice the electrolysis of water the boiling of water*

***Reactants*** are on the left side of the reaction arrow and ***products*** are on the right.

 *What are the names of the reactants in this reaction?*

 *HCl(aq) + NaOH(aq) 🡪 NaCl(aq) + H2O(l)* ***Reactants: aqueous hydrogen chloride (hydrochloric***

 ***acid) & aqueous sodium hydroxide***

***Coefficients*** are written in front of the formulas of reactants and products to balance chemical

equations. They give the ratios of reactants and products in a balanced chemical equation.

 *…****2****.Na + …….Cl2 🡪 …****2****NaCl*

*Only* **coefficients** can be changed when balancing chemical equations!

 *Given the unbalanced equation: Al + O2 🡪Al2O3*

*When this equation is balanced using the smallest whole numbers, what is the coefficient of Al?* ***4***

***Synthesis reactions*** occur when two or more reactants combine to form a single product.

*balance the following synthesis reaction*

***2****Na (s) + Cl2(g) 🡪* ***2****NaCl (s)*

***Decomposition reactions*** occur when a single reactant forms two or more products

 *balance the following decomposition reaction*

 *CaCO3(s) 🡪 CaO(s) + CO2 (g)*

***Single replacement reactions*** occur when one element replaces another element in a compound.

*Which equation below represents a reaction classified as a "single replacement" reaction?*

*2 H2 + O2 --> 2 H2O*

*Pb(NO3) 2 + K2CrO4 --> 2 KNO3 + PbCrO4*

 *HCl + NaOH --> H2O + NaCl*

 *Cu + Zn(NO3) 2 --> Zn + Cu(NO3) 2*

 *CaCO3 --> CO2 + CaO*

***Double replacement (displacement) reactions*** occur when two compounds react to form two new compounds.

 *Potassium sulfide (K2S) is mixed with lead acetate(Pb(C2H3O2). Which of the following products is*

 *expected? PbSO4 K2S K3PO4 PbS*

***Combustion reactions*** occur when a hydrocarbon burns in the presence of oxygen gas to produce carbon

 dioxide gas and water vapor. Write the combustion reaction for the combustion of methane gas (CH4)

 *CH4 (g) +* ***2O2 (g)*** *🡪* ***CO2 (g)*** *+* ***2H2O(g)*** *\*balance*

The masses (and energy) of the reactants in a chemical equation is always equal to the masses (and energy) of the products. “***Law of Conservation of Mass (and Energy)***.”

 *For the reaction: CaCO3 --> CO2 + CaO*

 *If 20.0g of CaCO3 decomposes to for 19.5 g of CaO, how many grams of CO­2 gas is formed?* ***0.5g***

The gram formula mass (**molar mass**) of a substance is the sum of the atomic masses of all the atoms in it. H2SO4 = **98.08** g/mole

 2 x H = 2 x **1.01** g = **2.02** g

 1 x S = 1 x **32.07** g = **32.06** g

 4 x O = 4 x **16.00** g = **64.00**g

Use the **mole** ***map*** to help you solve conversions

 What is the mass in grams of 3.6x1021 molecules of sugar? C12H22O11  = **2.05 g**

 What is the mass in grams of NaCl that was dissolved in 3.2 L to create a 1.6M solution? **299 g**

 An **empirical formula** is the simplest mole ratio among the elements in a compound.

 *Find the empirical formula of a compound composed of 75% carbon and 25% hydrogen.*

 ***75g/12.01 = 6.25 mol ÷ 6.25 = 1***

 ***25g/1.01 = 24.8 mol ÷ 6.25 = 4 so CH4***

A **molecular formula** is the true formula of a molecular compound.

 *Find the molecular formula of the compound from problem 32, if its molecular mass is 48.15 g/mol*

 **CH4  is 12.01g + 4(1.01g) = 16.05g/mol and 48.15 ÷ 16.05 = 3 so C3H12 is the molecular formula**

Know how to calculate the **percentage composition** of a compound.

  *Find the percent by mass of oxygen in CaCO3*.

**Molar Mass of CaCO3+ = 100.09 g/mol so %O = 3(16.00) ÷ 100.09 = 0.4795 x 100 = 47.95% oxygen**

Know how to convert an **empirical formula** into a **molecular formula**.

*A compound has the empirical formula NO2. Find its molecular formula if the molar mass = 92g.*

 *N2O NO2 N2O4* N*2*O

**NO2 = 14.01 + 2(16.00) = 46.01g/mol and 92 ÷ 46 = 2 so N2O4 is its molecular formula**

***Molarity*** is a way to measure the *concentration* of a solution.

Molarity is equal to the number of moles of solute divided by the number of liters of solution.

 *What is the molarity of an NaCl solution if 2.0 mol NaCl is present in 0.50 L solution?*

 **Molarity = 2.0mol ÷ 0.50L = 4.0 *M* solution**

Covalently bonded substances form **molecules**. Molecular compounds are made up of nonmetals only.

 *Which of the following are molecular compounds? H­2O CH4 NaCl CuSO4  CO2*

The formula of a **molecular compound** represents one molecule of the substance, also one mole of

 molecules and the number of atoms for each nonmetal element making up the molecule.

*The formula for methane gas, CH4, tells us the make-up of one molecule of methane is 1 carbon atom and 4 hydrogen atoms covalently bonded together. We can also figure out the molar mass or the mass in one mole of methane*. *What is the molar mass of CH4? 47 g/mol 24 g/mol 16 g/mol 10 g/mol*

What does it mean to say an element is **diatomic**? What are the 7 diatomic elements?

**Diatomics are elements who’s atoms are never found lone in nature, but in covalently bonded pairs.**

**Br2 I2 N2 Cl2 H2 O2 F2**

Write the balanced chemical equation for the following reaction. Be sure to include the symbols for the correct state of matter. Can you classify each reaction also?

*Crystal sugar (C12H22O11) is heated in the presence of oxygen gas to produce carbon dioxide gas and*

*water vapor.*

 **C12H22O11 (s) + 12 O2(g) Δ🡪 12CO2(g) + 11H2O(g)  Combustion**

*Aqueous magnesium sulfate reactions with aqueous barium chloride to produce solid barium sulfate and aqueous magnesium chloride.*

 **MgSO4(aq) + BaCl2(aq) 🡪 BaSO4(s) + MgCl2(aq) Double Replacement**