Math in Chemistry

1. Give a definition and example of a qualitative measurement. Give a definition and example of a quantitative measurement.

2. How many sig figs are contained in the following numbers:
   
   - 0.067
   - 2450
   - 2450.0
   - 0.0100

3. Define accuracy and precision.

4. If a block has a known density of 1.45, what is the percent error of the value 1.14 measured in lab?

5. A block of unknown composition has a mass of 23.2 grams. Using water displacement, its volume is measured at 27.9 mL. What is the block’s density?

6. What is the volume of an object with a density of 1.94 g/mL and a mass of 15.5g?

7. Convert the temperature 88°C to Kelvin.

8. Metric conversions: Convert 4.58 m to cm. Convert 104.5 J to kJ. Convert 0.009 332 g to mg.

9. What is the difference between a chemical change and a physical change?

Atomic Structure

An atom is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud). The atomic number is equal to the number of protons. The mass number is equal to the number of protons plus neutrons. In a neutral atom, the number of protons equals the number of electrons. The charge on an ion indicates an imbalance between protons and electrons. Too many electrons produces a negative charge, too few, a positive charge. This structure can be written as part of a chemical symbol.

Example:

\[
\begin{array}{c|c|c|c}
\text{isotope} & \text{atomic number} & \text{protons} & \text{neutrons} \\
^{13}\text{C} & 17 & 6 & 7 \\
^{17}\text{N} & 7 & 7 & 10 \\
^{26}\text{Ni} & 26 & 13 & 13 \\
^{37}\text{Ar} & 18 & 18 & 19 \\
^{52}\text{Fe} & 26 & 26 & 26 \\
^{50}\text{Ca} & 20 & 20 & 30
\end{array}
\]

1. Define the following terms:
   
   - atomic number
   - mass number
   - isotope
   - atomic mass

2. How do you determine the number of protons in an atom? How do you determine the number of neutrons in an atom?

3. Complete the following table:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic #</th>
<th># Protons</th>
<th># Neutrons</th>
<th>Mass #</th>
</tr>
</thead>
<tbody>
<tr>
<td>^{13}\text{C}</td>
<td>17</td>
<td>6</td>
<td>7</td>
<td>56</td>
</tr>
<tr>
<td>^{17}\text{N}</td>
<td>7</td>
<td>7</td>
<td>10</td>
<td></td>
</tr>
<tr>
<td>^{26}\text{Ni}</td>
<td>26</td>
<td>13</td>
<td>13</td>
<td></td>
</tr>
<tr>
<td>^{37}\text{Ar}</td>
<td>18</td>
<td>18</td>
<td>19</td>
<td></td>
</tr>
<tr>
<td>^{52}\text{Fe}</td>
<td>26</td>
<td>26</td>
<td>26</td>
<td></td>
</tr>
<tr>
<td>^{50}\text{Ca}</td>
<td>20</td>
<td>20</td>
<td>30</td>
<td></td>
</tr>
</tbody>
</table>

a. Consider the second and fourth row in the table above. What do they have in common?
b. Consider the second and fourth row in the table above. What are the differences?

**Isotopes and Average Atomic Mass**

Elements come in a variety of isotopes, meaning they are made up of atoms with the same atomic number but different atomic masses. These atoms differ in the number of neutrons. The average atomic mass is the weighted average of all the isotopes of an element.

**Example:** A sample of cesium is 75% $^{133}$Cs, 20% $^{132}$Cs and 5% $^{134}$Cs. What is its average atomic mass?

Answer: 
\[
\begin{align*}
.75 \times 133 &= 99.75 \\
.20 \times 132 &= 26.4 \\
.05 \times 134 &= 6.7 \\
\text{Total} &= 132.85 \text{ amu} = \text{average atomic mass}
\end{align*}
\]

Determine the average atomic mass of the following mixtures of isotopes.

1. 80% $^{127}$I, 17% $^{131}$I, 3% $^{128}$I

/ **Nomenclature Review:**

1. Metals form ions with ________________ charges, called _______________. Nonmetals form ions with ________________ charges, called _______________.
2. What is the difference between a molecular formula and a formula unit?
3. What region(s) of the periodic table are known as the representative elements?
4. How do you determine the charge of an ion from the representative elements?
5. Prefixes are used in the naming system for which type of compound (ionic or molecular)?
6. When naming ionic compounds, if the anion is a single atom, change the ending to ___________. If the anion is a polyatomic ion, use ________________. For molecular compounds, the number of atoms of each type is stated using ____________.
7. Ions are formed when atoms gain or lose what subatomic particle? ________________
Write the location on of each of the following families or classifications of elements on a periodic table: metals, nonmetals, metalloids, alkali metals, alkaline earth metals, halogens, noble gases, transition metals, & diatomic elements.

1. Horizontal rows are called ______________
2. Columns are called ___________ or ____________
3. Name group 1 _______________, 2 ________________, 3-12 ________________, 17 ________________, 18 ________________.
4. Elements are arranged according to their __________________
5. Elements within a group have the same number of __________________
6. Name the groups with the following ionic charges: +1 ____, +2 ____, +3 ____, -3 ____, -2 ____, -1 ___.
7. Label Zn and Ag with their charges.

---

The Periodic Table:

The letters inside the table have no significance here. They are from another worksheet.
8. As you go across the periodic table, the elements go from (metals / nonmetals) to (metal / nonmetals).


10. Where are the s, p, d and f sublevels?

11. Where are the most active metals?

12. Where are the most active nonmetals?

13. As you go across a period, the atomic size (decreases / increases).

14. As you go down a group, the atomic size (decreases / increases).

15. A negative ion is called a ___________ and is (larger / smaller) that its atom.

16. A positive ion is called a ___________ and is (larger / smaller) that its atom.

17. As you go down a group, the ionization energy generally (decreases / increases).

18. Where is the highest electronegativity found? Which element. Why?

19. Where is the lowest electronegativity found?

20. A colored ion generally indicates a ________________.

Complete the following table:

<table>
<thead>
<tr>
<th>Property</th>
<th>Trend → Period</th>
<th>Trend ↓ Group</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic Radius</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Electron Affinity</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Electronegativity</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ionization Energy</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Arrange the following atoms from largest to smallest atomic radius, and from highest to lowest ionization energy.

a. Cs, K, Li
b. Ba, Sr, Ca
c. I, Br, Cl
d. Mg, Si, S

1. Which has the larger atomic radius: O or F? Cl or Br? Ba or O?
2. Which has a higher electronegativity: O or F? Cl or Br? Ba or O?
3. Which has the lower ionization energy: O or F? Cl or Br? Ba or O?

Atomic Theory

1. Create an atomic theory timeline: List the following scientists/theories in chronological order. For all, draw a model of their atom and describe it.
   Bohr, Dalton, Quantum mechanical model, Rutherford, Thomson

   A. Dalton’s Theory
   1. All matter is composed of atoms.
   2. Atoms of the same element are the same and are different from atoms of other elements.
   3. Atoms cannot be subdivided, created, nor destroyed.
   4. Atoms combine in whole number ratios.
   5. Atoms in chemical reactions are combined, separated, or rearranged.
Electrons in Atoms

1. How do you write electron configurations?

2. Draw the orbital notation, electron configuration and noble gas configuration for: (Don't forget Hund's Rule!)
   - Iron, Oxygen, Nitrogen, Calcium, Chlorine, Bismuth, Tungsten

4. What is an orbital? Describe the shape of an s orbital. Describe a p orbital.

5. Write the correct electron configuration for the following elements:
   - Sulfur, Lead, Zirconium

6. How many unpaired electrons does sulfur contain?

7. Draw an electromagnetic wave. Label the crest, trough, wavelength, and amplitude.

8. Complete the following sentences:
   - As wavelength decreases, frequency _________________.
   - As wavelength decreases, energy _________________.
   - As frequency increases, energy _________________.

9. What are the two mathematical equations that relate wavelength, frequency and energy?

10. What is the wavelength of an electromagnetic wave that has a frequency of 65,200 Hz?

11. What is the energy of a photon of light with frequency $3.51 \times 10^{12}$ s$^{-1}$?

12. What type of radiation has the highest energy?

13. What type of visible light has the shortest wavelength?

14. When an electron absorbs a photon of light, it moves to a ____________ (higher/lower) energy level. What must the electron do to release this light energy?

15. In the Bohr model of the hydrogen atom, what electron transitions occur to produce the four lines of visible color in the visible spectrum? What other types of electromagnetic radiation are released by the Bohr hydrogen atom?
16. Explain what is meant by the wave-particle nature of light.
17. Are the colors of flame tests due to taking in energy or releasing energy? Explain.
18. What does it mean when we say energy levels are quantized?
19. What are the essential points of Bohr’s theory of the structure of the hydrogen atom?

Lewis dot diagrams – Draw the Lewis Dot structures for the elements listed below

Lewis diagrams are a way to indicate the number of valence electrons around an atom. Ca, Br, K, C, Ar, O, Al, He

Chemical bonding, Lewis Structures, Geometries
1. Define valence electron.
2. What is the octet rule?
3. What are resonance structures?
4. According to VSEPR Theory, what causes a water molecule to have a bent structure?
5. What makes a bond polar? What makes a molecule polar?
6. How can a molecule with polar bonds be a non-polar molecule overall?
7. Classify the following as ionic (metal/cation + nonmetal/anion), covalent (nonmetal and nonmetal) or both (compound containing a polyatomic ion).
   a) CaCl₂  b) CO₂  c) BaSO₄  d) K₂O  e) NH₄⁺  f) SiO₂  g) N₂  h) NO₃⁻  i) NH₄Cl  j) SO₃  k) NO₂  l) H₂O  m) CH₃Cl  n) H₂S
**Shapes of molecules**

The VSEPR theory helps us predict the shape of molecules. The unshared pairs of electrons will push the shared pairs away and change the shape of the molecule.

**DRAW** the Lewis dot structures for each of the above 15 structures and **DRAW** the dipole moments on each polar covalent bond. In polar covalent bonds the electrons are not shared equally. When the electronegativity difference is great enough (above 0.3) and the shape is asymmetrical, the molecule will be polar.

**Label the MOLECULAR GEOMETRY OF THE COVALENT MOLECULES** (ionic compounds are not molecules and do not have geometry).

**Then, label if the covalent molecules are likely to be polar or nonpolar according to their shape and polarity of bonds.**

**Which of the molecules in exhibit resonance?** Show the meaning of resonance with Lewis structures.

**Metallic Bonding:**

Describe the properties of metals and how their mobile electrons contribute to these properties. What is an alloy?

**Chemical Reactions**

1. In this equation: \(4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3\), what are the reactants? What are the products? What does the arrow mean?

2. Balance the following equation: \(\text{CuO} + \text{NH}_3 \rightarrow \text{Cu} + \text{H}_2\text{O} + \text{N}_2\)

3. What are the five types of chemical reactions?

4. Classify the following reactions:
   \[
   \begin{align*}
   \text{Mg} + \text{Zn(NO}_3\text{)_2} & \rightarrow \text{Mg(NO}_3\text{)_2} + \text{Zn} \\
   \text{H}_2 + \text{Cl}_2 & \rightarrow 2\text{HCl} \\
   \text{CH}_4 + 2\text{O}_2 & \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \\
   2\text{NaNO}_3 + \text{H}_2\text{SO}_4 & \rightarrow \text{Na}_2\text{SO}_4 + 2\text{HNO}_3 \\
   2\text{H}_2\text{O} & \rightarrow \text{H}_2 + \text{O}_2
   \end{align*}
   \]

5. Using the solubility rules in your reference tables, determine which compound is the precipitate in the following reaction:
   \(\text{Pb(NO}_3\text{)_2} + 2\text{KI} \rightarrow \text{PbI}_2 + 2\text{KNO}_3\)

6. Using the activity series in your reference tables, which of the following single replacement reactions will actually occur?
   \[
   \begin{align*}
   2\text{Na} + \text{FeSO}_4 & \rightarrow \text{Na}_2\text{SO}_4 + \text{Fe} \\
   3\text{CoCl}_2 + 2\text{Al} & \rightarrow 2\text{AlCl}_3 + 3\text{Co} \\
   2\text{KI} + \text{Cl}_2 & \rightarrow 2\text{KCl} + \text{I}_2
   \end{align*}
   \]

7. Predict the products of the following reactions. (HINT: determine what type of chemical reaction will occur and consult the solubility rules and activity series when necessary)
   \[
   \begin{align*}
   \text{AgNO}_3 + \text{CaCl}_2 & \rightarrow \text{K} + \text{ZnSO}_4 \\
   \end{align*}
   \]

8. List the 7 diatomic molecules ____________________________

9. Write the reaction for the following: Balance it. Determine if the reaction will occur by determining solubility. Write a net ionic equation and identify spectator ions.
• iron(III) chloride and sodium hydroxide
• silver nitrate reacts with potassium chloride
• aluminum chlorate reacts with magnesium phosphate
• Chromium(III) chloride solution plus fluorine gas
• Hydroiodic acid solution plus lead(II) nitrate solution
• Cadmium oxide decomposes when heated
• Nickel(II) sulfide solution plus sodium metal
• Acetic acid solution (HC2H3O2) is added to sodium hydroxide solution
• Magnesium metal reacts with oxygen gas
• Propane gas burns in oxygen gas
• Aluminum metal plus copper(II) iodide solution

**Composition Stoichiometry, The Mole, and Reaction Stoichiometry:**

**Percent Composition**

\[
\text{Percent Composition} = \left( \frac{\text{Mass of element}}{\text{Mass of the compound}} \right) \times 100\%
\]

Determine the percentage of potassium in potassium permanganate.

**Determining Empirical Formulas**

\[
\text{Percent Composition} \rightarrow \text{Emp Form}
\]

Determine the empirical if it is 75% carbon and 25% hydrogen.

**Determining Molecular Formulas**

The empirical formula of a compound is NO2. Its molecular mass is 92 g / mol. What is its molecular formula.

Problems

1. How many grams of C2H2 will be produced, if 7.00g of Ca(OH)2 are also produced in the following reaction?
   \[
   \text{CaC}_2 + 2\text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_2 + \text{Ca} (\text{OH})_2
   \]

2. How many liters of H2O will be produced, if \(3.60 \times 10^{23}\) molecules of CO2 are also produced in the following reaction at STP?
   \[
   \text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}
   \]

3. How many grams of H2O will be produced by the combustion of 58.7 liters of C2H4 at STP?
   \[
   \text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}
   \]

4. How many molecules of H2O are needed if 41.85g of O2 are to be produced from the reaction below?
   \[
   2\text{K}_2\text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{KOH} + \text{O}_2
   \]

5. How many grams of NO will be produced, if \(6.45 \times 10^{22}\) particles of H2O are also produced in the reaction below?
   \[
   4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}
   \]

6. In the chemical reaction below, how many grams of KCN will react with \(81.55\)g of H2SO4?
   \[
   2\text{KCN} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCN}
   \]

7. In the following unbalanced combustion reaction, how many liters of C8H18 will react with \(4.78\)g of O2 at STP?
   \[
   \text{C}_8\text{H}_{18} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
   \]

8. In this unbalanced chemical reaction \(\text{Fe}_3\text{O}_4 + \text{H}_2 \rightarrow \text{Fe} + \text{H}_2\text{O}\) How many grams of H2O will be produced, if \(41.95\)g of Fe3O4 are reacted completely with hydrogen?

9. Heating zinc sulfide in the presence of oxygen yields the following: \(2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2\)
a. If 1.72 mol of ZnS is heated in the presence of 3.04 mol of \( O_2 \), which reactant will be completely used up?
b. What mass of zinc oxide will be produced?

10. Calculate the percent yield for a reaction in which 0.38 g of \( \text{NO}_2 \) reacts and 0.36 g of \( \text{N}_2\text{O}_5 \) is recovered.

\[
2\text{NO}_2 + \text{O}_3 \rightarrow \text{N}_2\text{O}_5 + \text{O}_2
\]

10. Find the volume of 12.5 g of helium gas at STP.
11. Find the mass, in grams, of \( 2.03 \times 10^{24} \) molecules of \( \text{H}_2\text{O} \).
12. How many moles are in 174 g of \( \text{CaCl}_2 \)?
13. What is the percent composition of glucose (one type of sugar) whose formula is \( \text{C}_6\text{H}_{12}\text{O}_6 \)?
14. How many grams of silver nitrate must be added to an excess of sodium chloride to produce 42.3 g of silver chloride?

\[
\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3
\]

15. Calculate the number of grams of oxygen gas produced when 50.0 g of hydrogen peroxide decomposes according to the following UNBALANCED reaction: \( \text{H}_2\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{O}_2 \)

16. Determine the number of formula units of \( \text{KCl} \) produced when 25 g of \( \text{KClO}_3 \) react, given the following equation:

\[
2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2
\]

17. The UNBALANCED reaction of barium hydroxide with nitric acid is shown:

\[
\text{Ba(OH)}_2 + \text{HNO}_3 \rightarrow \text{H}_2\text{O} + \text{Ba(NO}_3)_2
\]

How many liters of water are produced when \( 3.85 \times 10^{21} \) particles of \( \text{Ba(OH)}_2 \) react with excess nitric acid?

18. If 155 g of \( \text{Pb(NO}_3)_2 \) react and 70.3 g of \( \text{NaNO}_3 \) are actually formed, what is the percent yield for the reaction:

\[
2\text{Na} + \text{Pb(NO}_3)_2 \rightarrow 2\text{NaNO}_3 + \text{Pb}
\]

Given the following UNBALANCED chemical reaction:

\[
\text{N}_2\text{O}_4(g) + \text{N}_2\text{H}_4(l) \rightarrow \text{N}_2(g) + \text{H}_2\text{O}(g)
\]

19. What is the limiting reagent when 5.00 g of nitrogen tetraoxide and 4.00 g of hydrazine (\( \text{N}_2\text{H}_4 \)) are mixed?

20. What mass of nitrogen is produced from the reaction described in question 8?
21. Determine the number of moles of 25 g of \( \text{NaCl} \).
22. Determine the number of grams of 2.5 mol \( \text{NaCl} \).
23. What volume will 3.2 moles of \( \text{O}_2 \) occupy at STP?
24. How many moles of hydrogen are needed to completely react with two moles of nitrogen?

\[
\text{N}_2 + 3\text{H}_2 \leftrightarrow 2\text{NH}_3 + 88\text{kJ}
\]

25. What volume of hydrogen is necessary to react with five liters of nitrogen to produce ammonia? (Assume constant temperature, pressure and they are all gases.)
27. How much heat would be produced if 6 moles of hydrogen gas reacted?
\[
N_2 + 3H_2 \leftrightarrow 2NH_3 + 88kJ
\]

**Limiting Reagents**

How many grams of ammonia can be produced from the reaction of 28 grams of nitrogen and 25 grams of hydrogen?  
\[
N_2 + 3H_2 \leftrightarrow 2NH_3 + 88kJ
\]

What is the limiting reactant?

If 30 g of ammonia is ACTUALLY produced, what is the percent yield?

**Kinetic Molecular Theory of Gases:** This is a description of IDEAL GASES that do not exist.

1. Gases consist of large numbers of tiny particles that are far apart.
2. Collisions are elastic.
3. Gas particles are in constant, rapid random motion.
4. There are no forces of attraction between gas particles.
5. The average kinetic energy of gas particles depends on temperature.

A. Which of the two from above are not completely true but accepted anyway?

An ideal gas will behave like a real gas when the TEMPERATURE IS LOW (slow moving particles are attracted to each other - hey baby) and the PRESSURE IS HIGH (particles can get close enough to bond - let's become a liquid. 😊).

Gases expand to fill their containers, they are fluid, have low density, are compressible, diffuse and effuse.

**Density**

\[ D = \frac{\text{mass}}{\text{Volume}} \]

**Gas Laws**

1. Define all the variables in the combined gas law and the ideal gas law.
2. Complete the following sentences:
   - When volume increases, pressure ________________.
   - When temperature increases, pressure ________________.
   - When temperature increases, volume ________________.
   - In gas law calculations, temperatures must always be in the unit ________________.
3. Complete the following calculations.
   HINTS: Decide which gas law to use first! Make sure units match at all times!
   a. A gas occupies 11.2 L at 0.860 atm. What is the pressure if the volume becomes 15.0L?
   b. A gas has a pressure of 699.0 mmHg at 40.0°C. What is the temperature at standard pressure?
   c. What volume would 32.0 g of NO₂ gas occupy at 3.12 atm and 18.0°C?
   d. Calculate the final pressure inside a scuba tank after it cools from 1.00 x 10³°C to 25.0°C.
      The initial pressure in the tank is 130.0 atm.
   e. A gas sample occupies 3.25 liters at 24.5°C and 1825 mmHg. Determine the temperature at which the gas will occupy 4250 mL at 1.50 atm.
4. Qualitatively explain Graham's Law.
In practical terms, it is often difficult to hold any of the variables constant. When there is a change in pressure, volume and temperature, the combined gas law is used.

\[
\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2} \quad \text{or} \quad P_1V_1T_2 = P_2V_2T_1
\]

Use the Ideal Gas Law below to solve the following problems.

\[
PV = nRT \quad \text{where} \quad P = \text{pressure in atmospheres} \\
V = \text{volume in liters} \\
n = \text{number of moles of gas} \\
R = \text{Universal Gas Constant} \\
0.0821 \text{ L atm/mol K} \\
T = \text{Kelvin temperature}
\]

1. How many moles of oxygen will occupy a volume of 2.5 liters at 1.2 atm and 25° C?

Dalton’s Law says that the sum of the individual pressures of all the gases that make up a mixture is equal to the total pressure or \( P_1 = P_1 + P_2 + P_3 + \ldots \). The partial pressure of each gas is equal to the mole fraction of each gas \( x \) total pressure.

\[
P_1 = P_1 + P_2 + P_3 + \ldots
\]

Solve the following problems.

1. A 250. mL sample of oxygen is collected over water at 25° C and 760.0 torr pressure. What is the pressure of the dry gas alone? (Vapor pressure of water at 25° C = 23.8 torr)

Graham’s Law says that a gas will effuse at a rate that is inversely proportional to the square root of its molecular mass, MM. Expressed mathematically:

\[
\frac{\text{rate}_1}{\text{rate}_2} = \sqrt{\frac{\text{MM}_2}{\text{MM}_1}}
\]

Solve the following problems.

1. Under the same conditions of temperature and pressure, how many times faster will hydrogen effuse compared to carbon dioxide?
### Heating and Cooling Curves

Change the temperatures to make these curves be for water.

1. What is a? b? c? d? e?
2. When does KE change?
3. When does PE change?
4. Which direction in endothermic? Exothermic?
5. Would the ΔT for an endothermic reaction be positive or negative?
6. What equation would solve for heat during lines a, c, e?
7. What equation would solve for heat during d? b?
8. What does it mean that the specific heat of water is more than the specific heat of ice?
9. What is the difference between temperature and heat?

### Thermochemistry:

1. Define all the variables in the following equations:
   - \( q = mC\Delta T \)
   - \( q = m\Delta H \)
2. Draw a heating curve and label the following:
   - solid, liquid, and gas region
   - melting and boiling
   - which equation from question two you would use on each segment of the curve
3. Complete the following calculations. HINT: decide whether it is a one-step or multi-step problem.
   a. How much heat is needed to raise the temperature of 7.9 g of iron from 22.6°C to 115°C. The specific heat if iron is 0.46 J/g °C.
   b. Calculate the amount of heat necessary to convert liquid water at 19.9°C to steam at 102.0°C.
4. Calculate the amount of heat evolved when 45.5 g of methane (CH\(_4\)) burns in an excess of air, according to the following equation:
   \[ CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H_{\text{rxn}} = -890.2 \text{ kJ} \]
Phase Diagram

1. What section represents the solid phase? Liquid? Gas?
2. Describe what line d-f represents.
3. Describe what line d-e represents.
4. Describe what line d-g represents.
5. What is the substance's normal melting point? Boiling Point?
6. When do all three phases exist at once?
7. Would an increase in pressure cause this substance to freeze or melt?

Solutions: Use the solubility curve below to answer questions 1-4.

1. How many grams of NaNO₃ will dissolve in 100mL of water at 30°C?
2. At 50°C, what gas has the greatest solubility in 100g of water?
3. What is the temperature at which 105 g of NaNO₃ will produce a saturated solution when dissolved in 100mL of water?
4. (a) At 40°C, what mass of KNO₃ is needed to produce a saturated solution?
   (b) At the same temperature, how would you describe a KNO₃ solution containing 57g of solute in 100g of water?

Terms to Know:
Soluble: capable of being dissolved
Insoluble: obviously not dissolvable
Saturated: contains maximum amount of dissolved solute
Supersaturated: a solution that contains more dissolved solute than a saturated solution.
Unsaturated: a solution that contains less solute than a saturated solution

1. What is the relationship between the solubility of gases and pressure?
2. What is the relationship between the solubility of gases and temperature?
3. Describe the type of solution, if 90 grams of sodium nitrate were dissolved in 100g 40C water. _______________ What about if the temp dropped to 10C? _____________

Molarity and Dilution

1. What is the concentration of a solution that has a volume to 450 mL and contains 201 grams of FeBr₂?
2. How many grams of K₂CO₃ are needed to make 200. mL of a 2.5 M solution?
3. What will the volume be if 100.5 grams of lithium bromide are used to make a 4.0 M solution?
4. What is the molarity of a solution that has a volume of 2.5L and contains 660g of Al₂(SO₄)₃?
5. A solution of 1.00 M NaCl is available. How many milliliters of this solution are needed to make a total of 100.0 mL of 0.750 M NaCl solution?

6. 2.00 L of 0.800 M NaNO₃ must be prepared from an available solution with a concentration of 1.50 M. What volume of the 1.50 M solution is required?

7. What is the molarity of a solution that has a volume of 500 mL and contains 25g of KCl?

8. How many grams of Mg(NO₃)₂ would you need to name 400 mL of a 1.5M solution?

9. How many liters of solution would you have if you used 80.0g of NaOH to make a 3.30M solution?

10. If 55.0mL of a 2.45M sucrose solution is diluted with water to make 175mL of solution, what is the final molarity?

**Compare and Contrast Table:**

<table>
<thead>
<tr>
<th>Solution</th>
<th>Colloid</th>
<th>Suspension</th>
</tr>
</thead>
<tbody>
<tr>
<td>Example</td>
<td>Salt water</td>
<td>Gel</td>
</tr>
<tr>
<td>Homogeneous vs. Heterogeneous</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Particle Size</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Settles when standing</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Can be filtered</td>
<td></td>
<td></td>
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<tr>
<td>Tyndall Effect</td>
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</tbody>
</table>

**Electrolytes**

Solutions with charges will conduct electricity and are therefore electrolytes. Electrolytes include:

1. Ionic solutes:
2. Strong Acids:
3. Strong Bases:

What is the difference between nonelectrolytes, weak electrolytes, and strong electrolytes? Give an example of each.

**Acids and Bases**

What are the three different definitions of acids and bases?

Acids: H⁺ / proton donors (electron pair acceptors)
Bases: H⁺ / proton acceptors (electron pair donors)

**STRONG ACIDS** become **weak conjugate bases**

<table>
<thead>
<tr>
<th></th>
<th>+ H₂O →</th>
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<tbody>
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<td>_____</td>
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</tbody>
</table>
Vapor Pressure and Boiling Point
1. When will a liquid boil?

2. What is the normal boiling point of A? B? C?
3. Which substance has stronger hydrogen bonds?
4. At higher altitudes, what will happen to the boiling points?

Changes to a System at Equilibrium
3 ways to apply a stress to a system
   a. Change the concentration
   b. Change the temperature
   c. Change the pressure
Le Chatelier's Principle states that when a system at equilibrium is subjected to a stress, the system will shift its equilibrium point in order to relieve the stress.

\[
12.6 \text{ kcal} + \text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI(}_\text{g})
\]

Which way will the system shift when:
1. Hydrogen is added?
2. Iodine is added?
3. Hydroiodic acid is added?
4. Iodine is removed?
5. Temperature is increased?
6. The pressure is increased?
7. When would pressure affect a system at equilibrium?
8. Write the equilibrium expression for this reaction.
9. What is K if [H_2] = 0.2, [I_2] = 0.4, and [HI] = 0.6

**Reaction Rate:**
The speed of the reaction OR The change in concentration of reactants over time.
What are the 5 factors affecting rate of reaction?
1. 2. 3. 4. 5.
\[\Delta E = \Delta H = \text{ENTHALPY = heat of the reaction = E}_\text{products} - E_\text{reactants}\]

Exothermic \(\Delta H\) = is negative
Endothermic \(\Delta H\) = is positive

1. Are the forward reactions endothermic or exothermic?
2. Which has more energy, reactants or products?
3. What does the catalyst do?
4. Is the catalyst part of the rxn?
5. Is \(\Delta H\) for this forward rxn positive or negative? How do you know?

**ENTROPY = ISDORDRE (Get it?)**
DISORDERLY = Gases > Solutions > Liquids > Solids = ORDERLY
DISORDERLY = 2 compounds > 1 compound = ORDERLY

**Facts about Reactions**
Indications that a reaction has taken place are
1. 2. 3. 4.

**Nuclear Decay**
Describe the characteristics

<table>
<thead>
<tr>
<th></th>
<th>symbols</th>
<th>masses</th>
<th>Shielding / penetrating ability</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alpha</td>
<td></td>
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<tr>
<td>Beta</td>
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<tr>
<td>Gamma</td>
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</tbody>
</table>
Predict the products of the nuclear reactions

Complete the following nuclear reactions and identify the reaction types:

a. $^{222}\text{Th} \rightarrow ^{210}\text{Po} + _____$

b. $^{235}\text{U} \rightarrow _0^0\text{e} + _____$

c. $^{105}\text{Ag} + _____ \rightarrow ^{105}\text{Pd}$

d. $^{11}\text{B} \rightarrow ^{11}\text{Be} + _____$

e. $^{234}\text{Th}^* \rightarrow ^{234}\text{Th} + _____$

4. Complete the following nuclear reactions:

a. $^{211}\text{Fr} \rightarrow _____ + ^{207}\text{Po}$

b. $^{82}\text{Kr} \rightarrow _____ + _____ \text{Kr}$

c. Technetium-99 undergoes beta decay

d. $^{237}\text{Np}$ releases an alpha particle

Concept of half-life:

1) Silicon-31 has a half-life of approximately 2.5 hours. If we begin with a sample containing 2000 kg of Si-31, what is the approximate amount remaining after 10 hours?

2) Carbon-15 has a half-life of 5.0 seconds. Suppose we have a sample containing 100 grams of carbon-15. How much will remain after 30 seconds?

3) Actinium-226 has a half-life of 29 hours. If 100. g of actinium-226 decays over a period of 87 hours, how many grams of actinium-226 remain?

Define Fission and define Fusion.... Make a drawing of each process.

How does a nuclear reactor work?