# 2015 Final Exam Information - Chemistry 821

**Date: Wednesday, June 17; 8:15 am- 10:00 am**

**Exam Room : Please check with your teacher for your exam room assignment**

**Book Return:**

**Time 7:40 –8:00 am on Wednesday, June 17 in room 3102**

**You may also return your book anytime during review week to your teacher**

**Remember to bring:**

* **#2 pencils**
* **It is your responsibility to bring in a NON PROGRAMABLE scientific calculator, no graphing calculators or cell phones – if you need to borrow a calculator from the department you can check it out during the book return. All borrowed calculators will have a 2% point penalty taken from your final exam grade.**
* **textbook**

**Weighting of final exam grade on your final grade in chemistry**

Option 1: The final exam counts as 12.5% and the midyear as 7.5% of your end of the year grade.

Option 2: The final exam counts as 20% and the midyear as 0% of your end of the year grade.

You will automatically get the higher of the two options.

Each term counts as 20% of your end of the year grade.

**Format:**

There will be a total of 80 multiple-choice questions, given in a 1 hour and 45 minutes block of exam time. The exam time includes the time it takes to pass out exam materials at the start of the exam.

You may use a non-programmable calculator. (No cell phones. No graphing calculators.) It is highly recommended that you do the practice problems with the calculator that you will use for the exam.

You will be given scrap paper, a periodic table with a formula sheet and a scantron sheet. You will need to bring a #2 pencil to record your answers on the scantron sheet correctly.

Each question is worth 1 point. You will get 1 point for every question answered correctly and 0 points for every question answered incorrectly or left blank.

**You should mark an answer for every question**.

If you have extended time as an accommodation you will be escorted to the extended time room at the end of the regular exam period. All exams must be finished and handed in before you are dismissed.

**Overview of topics covered on the exam**

Textbook: Zumdahl, Steven, et al. World of Chemistry. Illinois: McDougal Littell, 2007.

# Chapter Topic

1 Chemistry: an Introduction

1. Matter
2. Chemical Foundations: Elements, Atoms, and Ions
3. Nomenclature
4. Measurements and Calculations
5. Chemical Composition

7 Chemical Reactions: An Introduction

8 Reactions in Aqueous Solutions

9 Chemical Quantities

10 Energy

11 Modern Atomic Theory

1. Chemical Bonding

13 Gases

14 Liquids and Solids

1. Solutions
2. Acids and Bases
3. Equilibrium
4. 18 sec 1 & 2 Oxidation-Reduction Reactions
5. 19 sec. 1 Radioactivity and Nuclear Energy

**Study tips:**

* Organize and review all old exams and reading journals.
* Study sequentially.
* Divide your study time into short, intense sections. This can be more effective than studying continually for a long period of time.
* “Guess the test questions”. You should ask yourself what is most important when studying. What questions would you ask if you were writing the exam?
* Practice, practice, practice.
  + Go over the Review packet.
  + Go over suggested review problems in your textbook.
  + Do the standardized test prep questions at the end of each chapter.
  + Study with a friend. Quiz each other. Practice explaining topics to one another.
* For practice with multiple choice questions try the following:

1. On-line self-assessment quizzes for each chapter from the text book website.
2. Take the standardized test practice exams at the end of each chapter.
3. Take the practice exam.

**Suggested End of the Chapter Assessment Problems for final exam review**

Textbook: Zumdahl, Steven, et al. World of Chemistry. Illinois: McDougal Littell, 2007.

All blue numbered examples have answers in the back of the textbook.

# Chapter Topic

1. Chemistry: an Introduction   
   #14
2. Matter   
   #3, 5, 13, 16, 17, 22
3. Chemical Foundations: Elements, Atoms, and Ions   
   #18, 29, 32, 42, 52, 54
4. Nomenclature  
   #13, 21, 24, 29, 33
5. Measurements and Calculations  
   #7, 23, 24, 26, 34
6. Chemical Composition  
   #16, 22, 24, 29, 34, 36, 39, 50
7. Chemical Reactions: An Introduction  
   #26, 36
8. Reactions in Aqueous Solutions  
   #14, 18, 22, 27, 33
9. Chemical Quantities  
   #18, 28, 35, 40
10. Energy  
    #11, 14, 24, 28, 30, 33
11. Modern Atomic Theory  
    #28, 33, 37, 47, 54, 57
12. Chemical Bonding  
    #10, 34, 36, 40, 48, 49, 54

13 Gases  
#7, 14, 20, 27, 30, 34, 39, 45

14 Liquids and Solids  
 #15, 25, 28, 31, 36, 54

15 Solutions  
 #15, 19, 33, 38, 41  
16 Acids and Bases  
 #8, 25, 35

1. Equilibrium

#4, 5, 8, 32

18 sec 1 &2 Oxidation-Reduction Reactions and Electrochemistry  
#11, 19

19 sec. 1 Radioactivity and Nuclear Energy   
#13, 26

**Chapter 1: Chemistry and Chapter 2: Matter**

1. What is chemistry?

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

3. What are 5 evidences of a chemical change?

4. What is the Law of Conservation of Matter?

5. Arrange the following terms into a flow chart. Define each, and provide examples.

compound element

heterogeneous mixtures

matter made of atoms

covalent made of diatomic molecules

pure substances homogeneous

ionic made of formula units

6. Suggest one way a mixture can be separated into its elements.

7. Suggest one way a compound can be separated into its elements.

8. Define the following scientific terms:

a) Hypothesis

b) Scientific Theory

c) Scientific Law

9. How are intensive and extensive properties different? Give an example of each.

10. Label the following as either physical or chemical changes.

a) Slicing an apple \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

b) Boiling water \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

c) Two solutions mix and form a solid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

d) Cooking an egg \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

e) Breaking glass \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

11. Label the following as either a heterogeneous or homogeneous mixture, element, or compound.

a) peanut butter and jelly sandwich \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

b) water (H2O) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

c) copper \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

d) completely dissolved salt water \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

12, Complete this sentence: The mass of the reactants is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ equal to the mass of the products.

a) always b) sometimes c) never

**Chapter 19.1: Nuclear Chemistry and Radioactivity**

1. Define radioactivity? What type of nuclei tend to be radioactive?

2. Fill in the following table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Particle Type | Description | Isotope notation | Penetrating ability | Example Equation for decay |
| alpha |  |  |  |  |
| beta |  |  |  |  |
| gamma |  |  |  |  |

3. Write an equation for the beta decay of strontium-90

4. Write an equation for the alpha decay of uranium-238

5. Define half-life.

6. A given isotope has a half-life of 5.0 minutes. If the initial mass is 280 grams, how many grams will be left after 15 minutes? How many half-lives is this?

7. Write a balanced nuclear decay equation for each of the following:

1. Electron capture \_\_\_\_\_\_\_\_\_\_\_



1. Beta decay \_\_\_\_\_\_\_\_\_\_\_\_



1. Alpha decay \_\_\_\_\_\_\_\_



1. Positron Emitter \_\_\_\_\_\_\_\_\_\_\_\_\_



8. A substance has a mass of 2.50g after one half-life has occurred. What was the original mass?

9. Isotopes of the same element have the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ number of protons and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

number of neutrons.

**Chapter 3: Elements, Atoms, and Ions; Atomic Theory**

1. Compare the parts of an atom based on location, charge and mass:

- proton

- neutron

- electron

2. Define:

- isotope

- ion

- atomic number

- mass number

- atomic mass unit

3. How many neutrons does U-238 have?

4. Write isotope notation for the particle that contains 17 neutrons and 15 protons.

5. Fill in the following table:

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Name | Symbol | Atomic # | Mass # | # protons | # neutrons | # electrons |
| Sodium | Na | 11 |  |  | 11 |  |
| silver | Ag |  | 107 |  |  |  |
| copper (II) (cation) | Cu2+ |  |  | 29 | 31 |  |
| chloride (anion) | Cl- |  |  |  | 18 |  |
| Uranium 238(isotope) | 238U |  |  |  |  | 92 |

6. The atomic mass of carbon as displayed on the periodic table is 12.011 amu. However, no single carbon atom in nature has this mass. Explain.

7. If element Z (fictitious) has two isotopes: Z-20 (20.00 amu) with 91.2% abundance, and Z-21 (21.00 amu) with 8.8% abundance. If element Z were an actual element, what mass would be displayed on the periodic table?

8. Show the location of each of the following on the periodic table:

- periods

- groups or families

- main group elements

- metals

- non-metals

- metalloids

- alkali metals

- alkali earth metals

- transition metals

- halogens

- noble gases

- lanthanides

- actinides

9. Describe 4 properties of metals, and 4 properties of non-metals:

10. Which 2 elements touching the staircase are metals, and not metalloids?

11. Where are elements with similar properties found on the periodic table (in horizontal rows, or in vertical columns?)

12. Consider the element phosphorus.

1. period = \_\_\_\_\_\_\_\_\_
2. group = \_\_\_\_\_\_\_\_\_
3. atomic number = \_\_\_\_\_\_\_
4. atomic mass = \_\_\_\_\_\_\_\_\_ (round to the nearest whole number)
5. number of protons = \_\_\_\_\_\_\_\_
6. number of neutrons = \_\_\_\_\_\_\_\_
7. number of electrons = \_\_\_\_\_\_\_\_
8. Draw the Bohr diagram
9. Metal, or nonmetal? \_\_\_\_\_\_\_\_
10. What charge ion would phosphorus form?\_\_\_\_\_\_\_\_\_

13. What do elements lose or gain when ions are formed?

14. Describe how the periodic table is organized.

15. Elements in the same \_\_\_\_\_\_\_\_\_\_\_ tend to have similar properties because they have the same number of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

**Chapter 4: Chemical Formulas**

1. Name or write the formula for the following compounds:

a. NaBr \_\_\_\_\_\_\_\_\_\_\_ h. magnesium chloride \_\_\_\_\_\_\_

b. FeO \_\_\_\_\_\_\_\_\_\_\_ i. aluminum sulfate \_\_\_\_\_\_\_

c. Fe2(SO4­)3 \_\_\_\_\_\_\_\_\_\_\_ j. tin (II) chloride \_\_\_\_\_\_\_

d. Mg(NO3)2 \_\_\_\_\_\_\_\_\_\_\_ k. ammonium carbonate \_\_\_\_\_\_\_

e. KBr \_\_\_\_\_\_\_\_\_\_\_ l. sodium oxide \_\_\_\_\_\_\_

f. CaF2 \_\_\_\_\_\_\_\_\_\_\_ m. aluminum fluoride \_\_\_\_\_\_\_

g. FeCl3 \_\_\_\_\_\_\_\_\_\_\_ n. copper (II) fluoride \_\_\_\_\_\_\_

2. All of the above are a single type of compound. What type is it? \_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Name or write the formula for the following compounds:

a. carbon tetrachloride \_\_\_\_\_\_\_\_\_\_\_ e. Cl2O \_\_\_\_\_\_\_

b. boron trichloride \_\_\_\_\_\_\_\_\_\_\_ f. SO3 \_\_\_\_\_\_\_

c. dichlorine heptaoxide \_\_\_\_\_\_\_\_\_\_\_ g. P2O5 \_\_\_\_\_\_\_

4. All of the above are a single type of compound. What type is it? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

5. A metal and a nonmetal will form a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ bond, where electrons are exchanged between atoms.

6. A nonmetal and a nonmetal will form a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ bond, where electrons are shared between atoms.

7. Describe how you can recognize, from the formulas or names, each of the following:

- ionic compound

- covalent compound

- acids

- hydrate

8. For the following Compounds, write the chemical formula and circle either (I) for ionic or (M) for molecular or (A) for acid.

a) Sodium phosphide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

b) Aluminum oxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

c) carbon tetrachloride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

d) postassium hydroxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

e) Ammonium chloride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

f) calcium carbonate \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

g) carbonic acid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

h) phosphorous acid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

i) hydrochloric acid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

9. Write the names for each chemical formula and circle either (I) for ionic or (M) for molecular or (A) for acid.

1. NaClO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
2. BaS \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
3. Cl2O \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
4. SO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
5. HNO2 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
6. HF \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
7. SnCl2 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
8. Fe2 (CO3) 3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A
9. H2SO4 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ I M A

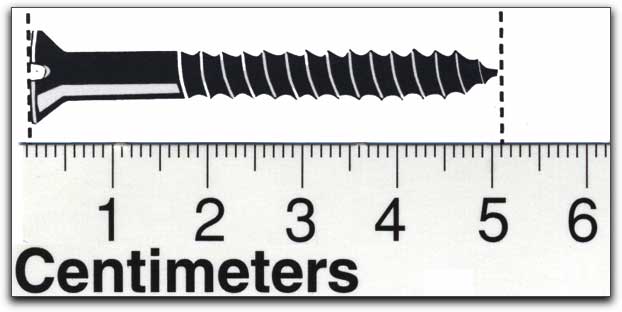
10. a. Draw the dot diagram for the ionic compound magnesium phosphide.

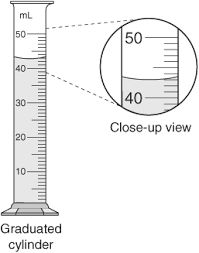
b. Write the chemical formula for the ionic compound magnesium phosphide.

**Chapter 5: Measurements and Calculations**

Scientific Figures/Measurements/Conversions

1. Measure the following to the correct number of significant figures:

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. How many significant digits does each of the following have?

a. 2300 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

b. 20040 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

c. 260.00 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

d. 0.00205 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

e. 4.65 x 10-4 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Answer the following with the correct number of significant digits.

a. 4.535 m + 0.0251 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

b. 274 m - 254 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

c. 6.54 m / 3.4215 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

d. 30.67 m x 23 m \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4. How many digits should be estimated in a measurement with a correct number of significant figures?

Temperature Scales

5. What is temperature? What is heat? Is heat the same as temperature?

6. Label these temperatures: Kelvin Celsius Fahrenheit

H2O freezing point \_\_\_\_\_ \_\_\_\_\_\_ \_\_\_\_\_\_

H2O boiling point \_\_\_\_\_ \_\_\_\_\_\_ \_\_\_\_\_\_

absolute zero \_\_\_\_\_ \_\_\_\_\_\_ \_\_\_\_\_\_

**Chapter 6: Chemical Composition (Calculations Involving the Mole)**

1. Distinguish between the following:

- empirical formula

- molecular formula

- structural formula

- condensed structural formula

2. A compound is 35.0% nitrogen, 5.0% hydrogen, and 60.0% oxygen. What is the empirical formula of the compound?

3. What is the mole?

4. How many atoms are in a mole of calcium atoms?

5. What is the mass of a mole of calcium atoms?

6. What is the mass of a mole of Mg(OH)2?

7. What is the percentage of silver in silver sulfide, Ag2S?

8. How many atoms are in 10.0 grams of aluminum?

9. How many grams is 3.4 x 1024 carbon atoms?

10. How much copper can be purified from 750 grams of copper (I) sulfide?

12. How many molecules of water are in 0.15 moles of H2O?

13. How many moles of nitrogen dioxide are in a sample that contains 5 x 1020 molecules NO2?

14. Calculate the number of moles in 9.0 g of H2O.

15. How many moles is 50 g of Ca3(PO4)2?

16. What is the volume of 15 mol of helium, He, at STP?

17. How many moles of oxygen, O2, are in 99 L of oxygen at STP?

18. a. What is the empirical formula of a substance that is 43.6% phosphorus and 56.4% oxygen by mass?

b. If the substance in part has a molar mass of 284 g/mole, what is its molecular formula?

19. How many oxygen atoms are in 10 molecules of Al2(SO4)3?

20. Calculate the percent composition of SO3. (Determine the %S and the %O.)

**Chapter 7: Types of Chemical Reactions and**

**Chapter 8: Reactions in Aqueous Solution**

1. In a chemical equation, what are the:

- reactants?

- products?

- subscript?

- coefficient?

- oxidation number?

symbols: (s) → (yield sign)

(l) ↑ (up arrow)

(g) ↓ (down arrow)

(aq) ∆ (delta sign)

(cr)

2. When balancing equations, which can you change: the subscripts, or the coefficients?

Why?

3. Balance the reaction: \_\_\_\_O2 + \_\_\_\_\_H2 🡪 \_\_\_\_\_H2O

4. Balance the reaction: \_\_\_\_\_Li + \_\_\_\_\_Ca3N2 🡪 \_\_\_\_\_Li3N + \_\_\_\_\_Ca

5. Describe what happens in each type of reaction. Give an example of each.

a. synthesis (combination)

b. decomposition

c. single replacement

d. precipitation (double replacement)

e. acid-base neutralization

f. combustion of a hydrocarbon

6. What are the 4 driving forces in a reaction in aqueous solution?

7. Using the solubility rules, tell which of these compounds are soluble:

**(S) = Soluble or (I) = Insoluble in water**

\_\_\_\_\_ a. magnesium hydroxide \_\_\_\_\_\_g. (NH4)3PO4

\_\_\_\_\_ b. silver chloride \_\_\_\_\_\_h. Al2S3

\_\_\_\_\_ c. barium sulfate \_\_\_\_\_\_ i. HgSO4

\_\_\_\_\_ d. potassium nitrate \_\_\_\_\_\_ j. Fe(OH)3

\_\_\_\_\_ e. lead (II) nitrate \_\_\_\_\_\_ k. CaCO3

\_\_\_\_\_ f. sodium carbonate \_\_\_\_\_\_ l. Co(NO­­­3)3

8. Balance, then Identify the Type of Reaction:

A = acid/base reaction

B = precipitation reaction

C = oxidation/reduction reaction

C1 = synthesis, C2 = decomposition, C3 = single replacement, C4 = combustion

\_\_\_\_ a. \_\_CdCO3(s) ------------> \_\_CdO(s) + \_\_CO2(g)

\_\_\_\_ b. \_\_Mg(s) + \_\_HCl(aq) --------> \_\_H2(g) + \_\_MgCl2(aq) ­

\_\_\_\_ c. \_\_CaBr2(aq) + \_\_AgNO3(aq) -----> \_\_Ca(NO3)2(aq) + \_\_AgBr(s)

\_\_\_\_ d. \_\_HCl(aq) + \_\_\_NaOH(aq) ----------> \_\_\_H2O(l) + \_\_\_NaCl(aq)

\_\_\_\_ e. \_\_PbCl2(aq) + \_\_Li2SO4(aq) --------> \_\_LiCl(aq) + \_\_PbSO4(s)

\_\_\_\_ f. \_\_As(s) + \_\_O2(g) --------> \_\_As2O5(s)

\_\_\_\_ g. \_\_CH4(g) + \_\_O2(g) ---------> \_\_CO2(g) + \_\_H2O(g)

\_\_\_\_\_ h. \_\_\_\_\_Ca + \_\_\_\_\_N2 🡪 \_\_\_\_\_Ca3N2

\_\_\_\_\_ i. \_\_\_\_\_Mg + \_\_\_\_\_ HCl 🡪 \_\_\_\_\_ MgCl2 + \_\_\_\_\_H2

\_\_\_\_\_ k. \_\_\_\_\_Pb(NO3)2 + \_\_\_\_\_NaI 🡪 \_\_\_\_\_ PbI2 + \_\_\_\_\_NaNO3

9. Write the word equation for each reaction in #8.

a.

b.

c.

10. Draw a molecular view of the reactions in #8 below:

g.



i.



10. Write the products and balance the following:

a. (synthesis) \_\_\_\_\_Li + \_\_\_\_\_O2 🡪

b. (decomposition)\_\_\_\_\_BaO 🡪

c. (precipitation) \_\_\_\_\_Ba(NO3)2 + \_\_\_\_\_H3PO4 🡪

d. (combustion) \_\_\_\_\_C3H8 + \_\_\_\_\_O2 🡪

e. (single replacement) \_\_\_\_\_Ca + \_\_\_\_\_HCl 🡪

f. (acid-base neutralization) \_\_\_\_\_KOH + \_\_\_\_\_HCl 🡪

11. Write the molecular, ionic, and net ionic equations for the reaction between:

barium chloride and sodium sulfate

**Chapter 9: Stoichiometry**

1. What is a mole ratio? Where in a balanced equation would you find the mole ratio?

2. Define:

- limiting reactant

- excess reactant

- theoretical yield

- percent yield

- percent error

**3. Mass to Moles**

Methane (CH4) reacts with oxygen gas to produce

Carbon dioxide and water:

\_\_\_CH4(g) + \_\_\_O2(g) --🡪 \_\_\_CO2(g) + \_\_\_H2O(g)­

5.0 grams of methane burns in air.

a. How many grams of water are made?

b. How many grams of carbon dioxide are made?

c. How many liters of carbon dioxide would this be equal to at STP?

**4. Moles to Moles**

Iron (II) oxide reacts with oxygen to form Iron (III) oxide:

\_\_\_\_FeO(s) + \_\_\_O2(g) --🡪 \_\_\_Fe2O3(s)­

a. Balance the equation with correct coefficients. What type of reaction is this?

b. How many moles of oxygen gas are required to react with 2.4 moles of iron (II) oxide?

c. How many moles of iron (III) oxide will be produced when 9.2 grams of iron (II) oxide react with excess oxygen?

**5. Limiting and Excess reactants**

Lithium metal combines with nitrogen gas to produce lithium nitride solid:

\_\_\_Li(s) + \_\_\_N2(g) --🡪 \_\_\_Li3N(s)­

a. Balance the equation with correct coefficients. What type of reaction is this?

b. If 56.0 grams of lithium are mixed with 56.0 grams of nitrogen gas, calculate the mass of lithium nitride produced from this reaction.

c. Which is the limiting reactant?

d. Which is the excess reactant? How many moles and grams are excess?

**6. Calculate the Percent Yield**

Methanol is an alcohol that is produced by the reaction of carbon monoxide and hydrogen gas (already balanced):

CO(g) + 2H2(g) --🡪­ CH3OH(l)

a. If If 6.5 x 104 grams of carbon monoxide are reacted with 8.5 x 103 grams of hydrogen, calculate the theoretical yield of methanol.

b. If 4.0 x 104 grams of methanol are actually produced, calculate the percent yield of methanol.

**7-13. Mixture of Stoichiometry questions**

7. Given the equation 2H2O🡪 2H2 + O2,

a. What is the ratio of moles of H2O to moles of O2?

b. How many moles of H2O would be required to produce 2.5 moles of O2?

8. Given the balanced equation 16HCl + 2KMnO4 🡪 2KCl + 2MnCl2 + 5Cl2 + 8H2O, if 1.0 mole of KMnO4 reacts, how many moles of H2O are produced?

9. Given the reaction CaCO3 🡪 CaO + CO2, if 3.00 moles of CaCO3 undergo decomposition to form CaO and CO2, how many grams of CO2 are produced?

10. In the reaction Zn + H2SO4 🡪 ZnSO4 + H2, how many grams of H2SO4 are required to produce 1.0 gram of H2?

11. Given the reaction CaCO3 🡪 CaO + CO2, if 50g CaCO3 react to produce 20g CO2, what is the percent yield of CO2?

12. Given the balanced equation: HCl + NaOH 🡪 H2O + NaCl, how many moles of NaCl are produced when 5 moles of NaOH are reacted?

13. Given the balanced equation: H2CO3 + 2NaOH 🡪 2H2O + Na2CO3 how many moles of sodium hydroxide are needed to produce 2.5 moles of Na2CO3?

**Chapter 10: Energy, Heat and Enthalpy**

1. Describe an exothermic and endothermic reaction in terms of energy of bond making and bond breaking.

2. Draw the energy diagram for an endothermic and exothermic reaction.

- Does the temperature increase or decrease in an endothermic reaction?

- Does the enthalpy of the system increase or decrease in an endothermic reaction?

3. For which type of reaction (endothermic or exothermic) is the sign of the enthalpy change negative? (Δ H < 0)

4. Label which process is exothermic or endothermic?

\_\_\_\_\_ a. when solid KBr (potassium bromide) is dissolved in water, the solution gets colder.

\_\_\_\_\_ b. natural gas CH4 (methane) is burned in a furnace.

\_\_\_\_\_ c. concentrated H2SO4 (sulfuric acid) is added to water, and the solution gets very hot.

\_\_\_\_\_ d. water boils in a teakettle.

5. How much energy is required to heat 7.40 mL of water from 25°C to 46°C?

(Assume the specific heat of water is 4.184 J/g°C)

6. If it takes 5.8 joules of energy to heat a piece of metal that weighs 1.6 grams from 23°C to 41°C, what is the specific heat of this metal? Is this metal pure gold (cAu­ = 0.13 J/g°C)? Why or why not?

7. The equation for the fermentation of glucose to alcohol and carbon dioxide is:

C6H12O6 ----------------------🡪 2 C2H5OH + 2 CO2 ∆H = - 67 kilojoules

a. Is this reaction endothermic or exothermic? Why?

b. Is energy released or absorbed?

c. How much heat is released when 25 moles of glucose is fermented?

8. Given the reaction

N2 (g) + O2 (g)🡨🡪 2NO (g) ∆H = +180.5 kJ

a. what will the enthalpy change be if 3 mol O2 reacts?

b. What will the enthalpy change be if 56 g N2 reacts?

**Chapter 11: Electron Configurations and Atomic Theory**

**Light, Photon Energies, and Atomic Spectra**

E = hv nano = 10-9

h = Planck’s constant = 6.63 x 10-34 J•s micro = 10-6

c = speed of light = 3.00 x 108 m/s milli = 10-3

1. What is light?

2. In the modern model of light, light behaves as a \_\_\_\_\_\_\_\_\_ and as a \_\_\_\_\_\_\_\_\_\_\_.

3. Calculate the wavelength of light that has a frequency of 3.20 x 1014 s-1

4. What is the frequency of electromagnetic radiation with a wavelength of 520 nm?

5. As the wavelength of light increases, the frequency \_\_\_\_\_\_\_\_\_ . Explain this relationship.

**Electromagnetic Spectrum**

6. What is electromagnetic radiation? Arrange the following from longest wavelength to the shortest wavelength:

visible light/gamma/radio/infrared/ultraviolet/microwaves

7. Which has the highest energy? Longest wavelength? Highest frequency?

**How Light Interacts with Matter**

8. How was Bohr’s model of the atom different from Rutherford's?

9. What is the difference between a continuous and a line spectrum?

10. Describe how energy is converted into light of specific wavelengths in the Bohr model of the hydrogen atom.

11. Explain why we see colors when an electron transitions from a higher energy state to a lower energy state. Include the terms, photon, excited state, and quantified energy.

12. Why do different elements have different line spectra?

13. How does an excited state electron configuration differ from a ground state electron configuration?

**Quantum Mechanics**

14. What is the Heisenberg uncertainty principle?

15. Describe the s and p orbitals in the quantum mechanical model.

16. How many orbitals are in the following sublevels?

3p \_\_\_\_\_ 2s \_\_\_\_\_ 4p \_\_\_\_\_

17. Arrange the following sublevels in order of decreasing energy: 2p, 4s, 3s, 3d, 3p

18. Identify the element that corresponds to the following electron configuration: 1s22s22p5 \_\_\_\_

19. Write the ground state electron configurations for

F \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Mg \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Fe \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Pb \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

O2 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Ca2+ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Fe2+ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

20. Draw an Aufbau diagram (orbital box diagram) for oxygen.

**Periodic Trends**

21. What is atomic radius? What region of the periodic table has the largest atomic radius? Smallest?

22. What is ionization energy? What region of the periodic table has the highest ionization energy? Lowest?

23. What is electronegativity? What region of the periodic table has the highest electronegativity? Lowest?

24. For the following elements – a. Ba, Ca, Ra b. P,Si,Al

- List the 3 elements in order from the largest atomic radius to the smallest atomic radius.

- List the 3 elements in order from the highest ionization energy to the lowest ionization energy.

**Chapter 12: Bonding and Molecular Geometry**

**12.1 - Characteristics of Chemical Bonds**

1. Describe the differences between these types of bonds:

- ionic bond

- covalent bond

- non-polar covalent bond

- polar covalent bond

2. What is a dipole? How do dipoles affect bonding?

3. How does the electronegativity value determine the polarity of a bond?

**12.2 - Characteristics of Ions and Ionic Compounds**

4. Which is generally larger than their parent atom - cations or anions? Why?

5. What is a polyatomic ion? How is a polyatomic ion different than simple cations or anions?

**12.3, 12.4 - Lewis Structures & VSEPR Theory**

6. Explain the "Octet Rule"

7. What is VSEPR Theory? Explain how VSEPR theory is used to predict the molecular geometry?

8. How do the bond angles in a tetrahedron, trigonal pyramid, and typical bent molecule compare? Why are they different?

9. a. Draw the correct Lewis Structures for the following.

b. Label the bonding electron pairs and the lone pair electrons.

c. Identify the bond angle and using VSEPR, predict the molecular geometry.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Molecule or Ion** | **Total valence electrons** | **Lewis Dot Structure** | **# Bonding domains**  **(on central atom)** | **# Lone pair electrons**  **(on central atom)** | **Bond Angle**  **Polar or non-polar?** | **Molecular Geometry** |
| O2 |  |  |  |  |  |  |
| NH3 |  |  |  |  |  |  |
| CH4 |  |  |  |  |  |  |
| H2O |  |  |  |  |  |  |
| BCl3 |  |  |  |  |  |  |
| CO |  |  |  |  |  |  |
| CO2 |  |  |  |  |  |  |

**Chapter 13: Gas Laws**

**13.1 - Kinetic-Molecular Theory of Gases and Gas Pressure**

1. Describe the Kinetic-Molecular Theory.

2. Describe and give units for each property of gases.

Pressure

Volume

Temperature

Moles

**13.2, 13.3 - Gas Laws and Avogadro's Hypothesis**

3. What is "STP"? Why is it used?

4. What is Avogadro's Hypothesis? What is the volume of 1 mole of any gas at STP?

5. Describe the relationship between pressure and volume.

6. Describe the relationship between volume and temperature in Kelvin.

7. Describe the relationship between pressure and temperature in Kelvin.

8. Describe the relationship between moles and volume.

9. Describe the relationship between moles and pressure.

10. As the temperature of a gas is raised, what happens to the average kinetic energy of the gas molecules? (increases, decreases, or stays constant)

12. Solve the following gas law problems. Be sure to state which gas law was used, and round all answers to the correct significant digits!

**State the Law and the numerical answer**

|  |  |
| --- | --- |
| \_\_\_\_\_\_\_\_\_\_  \_\_\_\_\_\_\_ moles | Calculate the number of moles of nitrogen gas are present if it occupies 25.3 Liters at 58.2 °C and 2.04 atmospheres of pressure? (R=0.0821 Latm/molK) |
| \_\_\_\_\_\_\_\_\_\_  \_\_\_\_\_\_\_ L | A balloon containing 4 liters of air at 45 °C is warmed to  90 °C. What is the new volume of the balloon if the pressure remains constant? |
| \_\_\_\_\_\_\_\_\_\_  \_\_\_\_\_\_\_ atm | A gas sample has a pressure of 1.4 atm and a volume of 5.5 L. What is the new pressure, if the gas volume increases to 11.0 L at constant temperature? |
| \_\_\_\_\_\_\_\_\_\_  \_\_\_\_\_\_\_ atm | You are scuba-diving in the Caribbean. The captain of the boat hands you a tank of pure compressed oxygen gas that has a volume of 15 liters at STP (1 atm, 0°C). What would the new pressure of the tank become if the temperature increases to 30 °C and volume decreases to 10 liters? |
| \_\_\_\_\_\_\_\_\_\_  \_\_\_\_\_\_\_ mmHg | The vapor pressure of water at 25 °C is 23.5 mmHg. If a sample of hydrogen is collected over water, find the partial pressure of hydrogen when the atmospheric pressure is 763 mmHg. |
| \_\_\_\_\_\_\_\_\_ atm | A sample of gas is transferred from a 175.0 mL vessel to a 50.0 mL vessel. If the initial pressure of the gas is 1.25 atm and if the temperature is held constant,   * 1. Will the pressure increase or decrease?   2. Calculate the pressure of this gas sample in the 50.0 mL container? |
| \_\_\_\_\_\_\_\_\_\_  \_\_\_\_\_\_\_ L | On a cold day, a person intakes 476 mL of air at 780 mm and –2.0 °C. What is the volume of this air in the lungs at 757 mm and 35°C? |
| \_\_\_\_\_\_\_\_\_ L | A balloon with a volume of 84 L at 25°C is taken outside where the temperature is 19°C.   1. Will the balloon’s volume get bigger or smaller? 2. Calculate the new volume. |
|  | 1. Calculate the number of moles of gas that has a volume of 84.2 L at 1.43 atm and 14°C. Given R = 0.0821 Latm/mol K |

**Chapter 14: Intermolecular Forces of Attraction and Phase Changes**

**14.1 - Intermolecular Forces of Attraction and Phase Changes**

1. What is the difference between:

- Intramolecular Bonding:

- Intermolecular Forces of Attraction:

2. Describe the 3 types of Intermolecular Forces of attraction:

- Dipole - Dipole:

- Hydrogen Bonding:

- London Dispersion Forces:

3. Describe and draw a picture hydrogen bonding between two water molecules.

4. Circle the correct statement about the properties of water.

a. high surface tension or low surface tension

b. high vapor pressure or low vapor pressure

c. polar or nonpolar

5. Explain why aqueous NaCl is conductive, but solid NaCl is not.

6. Explain why an aqueous solution of table salt (NaCl) contains electrolytes, but an aqueous solution of table sugar (C12H22O111) contains non electrolytes.

7. a. What is a hydrate?

b. Describe how the water of hydration could be removed from a hydrate.

8. Given the following data, determine the %water in the hydrate.

Mass of empty crucible = 15.00 g

Mass of hydrate & crucible before heating = 17.50 g

Mass of crucible & contents after heating = 16.50 g

9. What is a phase change? Describe these 6 phase changes :

1. melting =

2. freezing =

3. evaporation =

4. condensation =

5. deposition =

6. sublimation =

10. Label the phase change diagram for water with:

- solid phase - melting point - temperature (Celsius)

- liquid phase - boiling point - pressure (atm)

- gas phase - triple point - critical temp./pres.

11. Label the heating/cooling curve for water with:

- ice - ice and water - boiling point

- water - water and steam

- steam - melting point

**14.2 - Vapor Pressure and Boiling Point**

12. What is vapor pressure?

13. What is boiling point? Melting point?

**14.3 - Properties of Solids**

15. Which has the higher boiling point? Why?

a. NaCl or O2 ­

b. CO­2 or KBr

c. H2 or H2O

16. Match the type of solid with the correct description

\_\_\_\_ sucrose (table sugar) a. covalent network solid

\_\_\_\_ diamond b. ionic solid

\_\_\_\_ sodium chloride (table salt) c. molecular solid

\_\_\_\_ steel d. metal

\_\_\_\_ titanium e. alloy

**Chapter 15 & 16: Solutions and Acids & Bases**

**15.1 - Solution Properties**

1. Define:

- Solution

- Solute

- Solvent

- Saturated Solution

- Unsaturated Solution

- Supersaturated

- Dilute

- Concentrated

- Solubility

2. What are 3 Ways to Increase the Solubility of a substance?

**15.2, 15.3 - Solution Composition; Solution Stoichiometry**

3. Calculate the mass percent of 5.00 grams calcium chloride in 95.0 grams of water

4. Calculate the mass in grams of NaCl present in 11.5 grams of 6.25% NaCl solution

5. If 0.50 moles of KBr is present in 250 milliliters of solution, calculate the molarity of the solution.

6. If 45.3 grams of KNO3 is dissolved in enough water to make 225 mL of solution, what is the molarity?

7. Calculate the new molarity when 250 mL of water is added to 125 mL of 0.250 M

HCl solution to dilute the acid.

8. a. Write a balanced equation for the neutralization reaction between HCl and NaOH.

b. What volume of 1.0 M HCl is needed to neutralize 10 mL 2.0 M NaOH.

**16.1, 16.2 - Properties of Acids & Bases; pH Scale**

9. What is an acid? List 5 properties of acids.

1. 4.

2. 5.

3.

10. What is a base? List 5 properties of bases.

1. 4.

2. 5.

3.

11. What is Kw = \_\_\_\_\_\_\_\_\_\_ at 25°C?

12. What is "pH"? What is the formula for pH?

13. As the pH value increases, the [H+] hydronium ion concentration \_\_\_\_\_\_\_\_\_\_

14. For each of the following solutions, tell the pH value, the [H+], and the [OH-] .

pH range [H+] (high/low) [OH-] (high/low)

Acid solution \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

Base solution \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

Neutral solution \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

15. What is a conjugate acid? What is a conjugate base?

16. Given the following concentrations of hydrogen and hydroxide ions, find the pH

values and tell whether the solution is acidic or basic!

Concentration of ion pH Acid or Base

[H+] = 1 x 10-2 \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

[OH-] = 1 x 10-2 \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

[H+] = 1 x 10-11 ­­­­­­­\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

[OH-] = 1 x 10-8 \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

[H+] = 1 x 10-7 \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

17. Find the pH or the concentration of ion of these "decimal" acids and bases!

Concentration of ion pH Acid or base

[H+] = 2.4 x 10-6 \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

[H+] = 9.1 x 10-9 \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

[H+] = \_\_\_\_\_\_\_\_\_\_\_\_ 13.2 \_\_\_\_\_\_\_\_\_\_

**Chapter 17: Equilibrium and Kinetics**

1.Use the collision model to explain how a chemical reaction occurs. Be sure to include what is necessary for a collision to be successful in producing a reaction.

2. Explain how the following affect reaction rate:

a. temperature

b. concentration

c. surface area

d. catalyst

3. Describe a system at chemical equilibrium.

4. Write the equilibrium constant expression "K" for the reaction:

2 SO2(g) + O2(g) 🡨🡪 2 SO3(g)

a. N2(g) + 3 H2 (g) 🡨🡪 2NH3 (g)

b. Fe(OH) 3 (s) 🡨🡪 Fe 3+(aq) + OH- (aq)

c. H2O (l) 🡨🡪 H+(aq) + OH- (aq)

d. HNO2 (aq) 🡨🡪 H+ (aq) + NO2- (aq)

5. Given the following equation,

2 SO2 (g) + O2 (g)🡨🡪 2SO3 (g) ∆H < 0

predict the equilibrium shift caused by each of the following changes

1. SO3 is added \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. SO2 is removed \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
3. The volume is increased. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
4. The temperature is increased. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Chapter 18: Reduction/ Oxidation and Electrochemistry**

1. Define:

a. oxidation

b reduction

2. What happens to the oxidation number of an element in a compound undergoing oxidation?

3. What happens to the oxidation number of an element in a compound undergoing reduction?

LEO says GER

4. Write the oxidation numbers of each element in the following:

a. CaO

b. AlCl3

c. Fe2S3

d. O2

e. KMnO4

5. For the reaction below, identify the element that is reduced, the element that is oxidized, the reducing agent, the oxidizing agent, the reduction half-reaction, the oxidation-half reaction, and then balance the chemical equation by half-reactions showing a balance of electrons transferred.

\_\_\_\_ Zn (s) + \_\_\_\_ Ag+ (aq) 🡪 \_\_\_\_ Zn2+ (aq) + \_\_\_\_ Ag (s)

Element reduced: \_\_\_\_\_\_\_\_\_\_ Element oxidized: \_\_\_\_\_\_\_\_\_\_\_\_

Reducing agent: \_\_\_\_\_\_\_\_\_\_\_ Oxidizing agent: \_\_\_\_\_\_\_\_\_\_\_\_\_

Reduction half-reaction: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Oxidation half-reaction: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

6 a. Draw a standard voltaic cell, Ni| Ni2+  || Cu2+ +|Cu.

b. Label the anode, cathode, and salt bridge. An Ox – Red Cat

c. Draw in an arrow to indicate the direction of electron flow along the wire.

**Formula Sheet**

**Equations**

|  |  |
| --- | --- |
| P V = n R T  =    =  Kw = [H3O+] [OH-]  pH = - log [H3O+] | q = m c ∆T  M1V1= M2V2  Molarity M = #moles solute/ L solution  Molality m = #moles solute/ kg of solvent  TK = T°C + 273  c = ν λ  E = h ν |

**Constants**

|  |  |
| --- | --- |
| Gas law constant = R = 0.08206 L atm/mol K  R = 8.314 L kPa/mol K  R = 62.4 L mmHg/mol K  1 atm = 760 mmHg = 101.3 kPa  Avogadro’s constant = NA = 6.022 x 1023/mol  Speed of light = c = 2.998 x 108 m/s | Planck’s constant = h = 6.626 x 10-34 J s  Equilibrium constant for water = Kw = 1 x 10-14 at 25°C  Specific heat of water = 4.184 J/g°C |

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Common Polyatomic Ions** | | | | **General Rules for Solubility of Ionic Compounds (Salts) in Water at 25 ° C** |
| NH4+ | Ammonium | CO32- | Carbonate | 1. Most nitrates (NO3-) salts are soluble. 2. Most salts of Na+, K+, and NH4+ are soluble. 3. Most chloride salts are soluble. Notable exceptions are AgCl, PbCl2 and Hg2Cl2. 4. Most sulfate salts are soluble. Notable exceptions are BaSO4, PbSO4 and CaSO4. 5. Most hydroxide compounds are only slightly soluble.\* The important exceptions are NaOH and KOH. Ba(OH)2 and Ca(OH)2 are moderately soluble 6. Most sulfides (S2-), carbonates (CO32-) and phosphates (PO43-) salts are only slightly soluble.   \*The terms *insoluble* and *slightly soluble* really mean the same thing: such a tiny amount dissolves that it is not possible to detect with the naked eye. |
| NO2- | Nitrite | HCO3- | Hydrogen carbonate (bicarbonate) |
| NO3- | Nitrate |
| SO32- | Sulfite | ClO- | Hypochlorite |
| SO42- | Sulfate | ClO2- | Chlorite |
| HSO4- | Hydrogen sulfate (bisulfate) | ClO3- | Chlorate |
| ClO4- | Perchlorate |
| OH- | Hydroxide | C2H3O2- | Acetate |
| CN- | Cyanide | MnO4- | Permanganate |
| PO43- | Phosphate | Cr2O72- | Dichromate |
| HPO42- | Hydrogen phosphate | CrO42- | Chromate |
| H2PO4- | Dihydrogen phosphate | O22- | Peroxide |

**The Periodic Table of Elements**

**1A 8A**

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 1  **­H**  1.008 | **2A** |  |  |  |  |  |  |  |  |  |  | **3A** | **4A** | **5A** | **6A** | **7A** | 2  **He**  4.003 |
| 3  **Li**  6.941 | 4  **Be**  9.012 |  |  |  |  |  |  |  |  |  |  | 5  **B**  10.81 | 6  **C**  12.01 | 7  **N**  14.01 | 8  **O**  16.00 | 9  **F**  19.00 | 10  **Ne**  20.18 |
| 11  **Na**  22.99 | 12  **Mg**  24.31 |  |  |  |  |  |  |  |  |  |  | 13  **Al**  26.98 | 14  **Si**  28.09 | 15  **P**  30.97 | 16  **S**  32.07 | 17  **Cl**  35.45 | 18  **Ar**  39.95 |
| 19  **K**  39.10 | 20  **Ca**  40.08 | 21  **Sc**  44.96 | 22  **Ti**  47.88 | 23  **V**  50.94 | 24  **Cr**  52.00 | 25  **Mn**  54.94 | 26  **Fe**  55.85 | 27  **Co**  58.93 | 28  **Ni**  58.69 | 29  **Cu**  63.55 | 30  **Zn**  65.39 | 31  **Ga**  69.72 | 32  **Ge**  69.72 | 33  **As**  74.92 | 34  **Se**  78.96 | 35  **Br**  79.90 | 36  **Kr**  83.80 |
| 37  **Rb**  85.47 | 38  **Sr**  87.62 | 39  **Y**  88.91 | 40  **Zr**  91.22 | 41  **Nb**  92.91 | 42  **Mo**  95.94 | 43  **Tc**  (98) | 44  **Ru**  101.1 | 45  **Rh**  102.9 | 46  **Pd**  106.4 | 47  **Ag**  107.9 | 48  **Cd**  112.4 | 49  **In**  114.8 | 50  **Sn**  118.7 | 51  **Sb**  121.8 | 52  **Te**  127.6 | 53  **I**  126.9 | 54  **Xe**  131.3 |
| 55  **Cs**  132.9 | 56  **Ba**  137.3 | 57  **La**  138.9 | 72  **Hf**  178.5 | 73  **Ta**  181.0 | 74  **W**  183.8 | 75  **Re**  186.2 | 76  **Os**  190.2 | 77  **Ir**  192.2 | 78  **Pt**  195.1 | 79  **Au**  197.0 | 80  **Hg**  200.6 | 81  **Tl**  204.4 | 82  **Pb**  207.2 | 83  **Bi**  209.2 | 84  **Po**  (209) | 85  **At**  (210) | 86  **Ru**  (222) |
| 87  **Fr**  (223) | 88  **Ra**  226.0 | 89  **Ac**  227.0 | 104  **Rf**  (261) | 105  **Db**  (262) | 106  **Sg**  (263) | 107  **Bh**  (262) | 108  **Hs**  (265) | 109  **Mt**  (266) | 110  (269) | 111  (272) | 112  (277) |  | 114  (289) |  | | | |
|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  | | | |
|  |  | 58  **Ce**  140.1 | 59  **Pr**  140.9 | 60  **Nd**  144.2 | 61  **Pm**  (145) | 62  **Sm**  150.4 | 63  **Eu**  152.0 | 64  **Gd**  157.3 | 65  **Tb**  158.9 | 66  **Dy**  162.5 | 67  **Ho**  164.9 | 68  **Er**  167.3 | 69  **Tm**  168.9 | 70  **Yb**  173.0 | 71  **Lu**  175.0 |  |  |
|  |  | 90  **Th**  232.0 | 91  **Pa**  231.0 | 92  **U**  238.0 | 93  **Np**  237.0 | 94  **Pu**  (244) | 95  **Am**  (243) | 96  **Cm**  (247) | 97  **Bk**  (247) | 98  **Cf**  (251) | 99  **Es**  (252) | 100  **Fm**  (257) | 101  **Md**  (258) | 102  **No**  (259) | 103  **Lr**  (260) |  |  |