Worksheet 1: Chapter 1 Lecture Problems

1. Write the following in scientific notation:  
   - 345.9 nm
   - 0.0054 mL
   - 0.00000050 kg
   - 93,000,000 miles (ave. distance from Earth to Sun)
   - 0.0000000000000000000000167 g

2. Calculate: \(15.05 \times [1.6 \times 10^{-3}] \div [9.46 \times 10^{-8}] = \)

3. Convert:  
   - 45 m to cm
   - 35.8 \(\mu\)s to cs
   - 0.012 km to m
   - 0.00095 km to nm

4. How many cm\(^2\) in 1 m\(^2\)?  
   - How many cm\(^2\) are in an object 3.94 in \(\times\) 3.94 in?
   - How many cm\(^3\) are in 1 m\(^3\)?  
   - How many cc are in a 125 in\(^3\) engine?

5.  
   | How many sig figs in each number | 
   |-----------------|----------------|
   | 3.4680           | 34,608         |
   | 0.3468           | 3,460,800      |
   | 0.03468          | 3.4600 \(\times\) 10\(^{-2}\) |
   | 0.034680         | 3.46 \(\times\) 10\(^{-2}\) |

6.  
<table>
<thead>
<tr>
<th>Underline the sig figs</th>
</tr>
</thead>
<tbody>
<tr>
<td>303.68 m</td>
</tr>
<tr>
<td>0.010 sec</td>
</tr>
<tr>
<td>5.883(\times)10(^{12}) atoms</td>
</tr>
</tbody>
</table>

7. Rules for Sig Figs:  
   - For multiply/divide, use the _____________________________
   - For addition/subtraction, use the _____________________________

8. \(\frac{(1.01\times10^7) \times (4.059\times10^5)}{(319.4 \ 318.7) \times (4.2 \ 9.7)}\)

9. \((3.052 \ km \ + \ 1114.2 \ m) \div 62.12 \ s\)

10. The density of iron is 7.86 g/cm\(^3\). What is the volume of 10.5 g of iron?

11. Convert the following temperatures  
    - a. 83.2 °C to °F  
    - b. 15.8 °F to K
Worksheet 2: Significant Figures and Conversions

Extra Problems of this type at end of course packet!

1. Count the number of significant figures in the following numbers:
   a. 303.68        b. 0.010        c. 5.883 \times 10^{32}

2. Write the numbers from problem #1 a & b in scientific notation.

3. Perform the following calculations and give the answer with the correct number of significant figures.
   a. (43.943)(8.98)
   b. (6.857 \times 10^5)(4.1 \times 10^3)
   c. 87.255 + 944.3
   d. \frac{176.158 - 17.2}{3789.55}
   e. (543.1)(16.8) + 326.978

4. Fill in the blanks.
   a. _____ \mu g = ___1___ g        b. _____ s = ___1___ ps

5. Perform the following conversions. Use the conversion chart in your textbook (but don’t peek at the metric prefixes). Don’t forget significant figures!
   a. 347 \mu g to g
   b. 6.724 \times 10^{-5} \text{ km to mm}
   c. 764.6 \text{ lb to kg}
   d. 3.6 \text{ gal to mL}
   e. 56.2 \text{ in}^2 to \text{ cm}^2

6. What is the density of a liquid if 156 mL has a mass of 178.2 g?

7. The density of iron is 7.86 g/cm^3. What is the volume of 10.5 g of iron?

8. Convert the following temperatures
   a. 83.2 °C to °F        b. 15.8 °F to K
Worksheet 2 cont. Conversions and Significant Figures and Moles

Convert:

1. \(1.487 \times 10^{-8}\) nm to m
2. \(4.15 \times 10^7\) pm to \(\mu m\)
3. \(8.97 \times 10^{-4}\) MW to mWatts
4. \(0.000429\) km\(^3\) to cm\(^3\)
5. \(59.42\) ft\(^2\) to m\(^2\)
6. \(6.47 \times 10^{-2}\) L to mm\(^3\)
7. \(19.5\) g/cm\(^3\) to kg/m\(^3\)
8. \(10.5\) g/mL to pound/gallon
9. \(8.29\)¢/pound to $/kg
10. \(1\) km/100 L to miles/gallon

Answer with the proper Significant figures:

11. \((1.009 \times 10^{-2} \text{ m} + 2.914 \times 10^{-3} \text{ m}) \times 4.129 \times 10^{-9} \text{ m} =\)
12. \(9.2735 / (9.457 - 8.693) =\)
13. \((4.598 + 33.7 + 57.1) \times 2.9178 \times 10^{-4} =\)
14. \((1.41 \times 10^{-7} \times 2.98 \times 10^{-5} \times 1.10 \times 10^2) \div (4.129 \times 10^{-3})^3\)

Mass–Mole ratios

15. Calculate the molar mass of dinitrogen tetraiodide, phosphorus pentachloride, and aluminum phosphate.

16. How many mol of Ne is in 89.7 g of this gas? [Ne has no known compounds; it seems to be inert.]

17. How many mol of Hg atoms are in 1.63 \(\times\) 10\(^{-2}\) g of the metal? [Hg is liquid at room temperature]

18. How many oxygen atoms (in an actual number) are in 7.42 \(\times\) 10\(^{-14}\) moles of ozone [O\(_3\)] molecules?

19. What mass, in g, of lead is present in 5.29 \(\times\) 10\(^{-3}\) mol of the metal?

20. What mass, in mg, of sodium sulfate is present in 4.51 \(\times\) 10\(^{-4}\) mol of this compound?

How many moles of ammonium carbonate corresponds to 8.715 g of this compound?
# Worksheet 3: Name That Compound!

Please either give the proper name or write the correct chemical formula to fill in the table.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium oxide</td>
<td>CrCl₂</td>
</tr>
<tr>
<td>calcium oxide</td>
<td>Co₂(SO₃)₃</td>
</tr>
<tr>
<td>aluminum bromide</td>
<td>LiOH</td>
</tr>
<tr>
<td>aluminum oxide</td>
<td>BCl₃</td>
</tr>
<tr>
<td>magnesium carbide</td>
<td>S₂O₇</td>
</tr>
<tr>
<td>chromium (II) sulfate</td>
<td>Pb(NO₃)₂</td>
</tr>
<tr>
<td>chromium (III) sulfide</td>
<td>Pb(C₂H₃O₂)₄</td>
</tr>
<tr>
<td>manganese (III) hypochlorite</td>
<td>I₂O₄</td>
</tr>
<tr>
<td>sulfur trioxide</td>
<td>NaHCO₃</td>
</tr>
<tr>
<td>diphosphorous pentoxide</td>
<td>HCl (g)</td>
</tr>
<tr>
<td>ammonium carbonate</td>
<td>HCl (aq)</td>
</tr>
<tr>
<td>zinc cyanide</td>
<td>K₂CrO₄</td>
</tr>
<tr>
<td>silver phosphate</td>
<td>FeI₃</td>
</tr>
<tr>
<td>dihydrogen monoxide</td>
<td>CuSO₄</td>
</tr>
<tr>
<td>potassium chlorate</td>
<td>SiF₄</td>
</tr>
<tr>
<td>magnesium nitrite</td>
<td>Ca(NO₃)₂</td>
</tr>
<tr>
<td>iron (II) acetate</td>
<td>Pb(CO₃)₂</td>
</tr>
<tr>
<td>nitric acid</td>
<td>Co₃(PO₄)₂</td>
</tr>
<tr>
<td>nitrous acid</td>
<td>H₂SO₄</td>
</tr>
</tbody>
</table>
Worksheet 4: Balancing Equations

1) _____ Co(s) + _____ O_2(g) → _____ Co_2O_3(s)

2) _____ LiClO_3(s) → _____ LiCl(s) + _____ O_2(g)

3) _____ Cu(s) + _____ AgC_2H_3O_2(aq) → _____ Cu(C_2H_3O_2)_2(aq) + _____ Ag(s)

4) _____ C_2H_11OH(l) + _____ O_2(g) → _____ CO_2(g) + _____ H_2O(g)

5) _____ H_2SO_4(aq) + _____ Al(OH)_3(aq) → _____ Al_2(SO_4)_3(aq) + _____ H_2O(l)

6) _____ H_2CO_3(aq) + _____ NH_4OH(aq) → _____ (NH_4)_2CO_3(aq) + _____ H_2O(l)

7) _____ C_6H_12O_6(s) + _____ O_2(g) → _____ CO_2(g) + _____ H_2O(g)

8) _____ LiNO_3(s) → _____ LiNO_2(s) + _____ O_2(g)

9) _____ Pb(s) + _____ O_2(g) → _____ PbO(s)

10) _____ NaOH(aq) + _____ HCl(aq) → _____ NaCl(aq) + _____ H_2O(l)

11) _____ PCl_5(s) + _____ H_2O(l) → _____ H_3PO_4(aq) + _____ HCl(aq)

12) _____ SiCl_4(s) + _____ H_2O(l) → _____ SiO_2(s) + _____ HCl(aq)

13) _____ CaO(s) + _____ P_4O_10(l) → _____ Ca_3(PO_4)_2(s)

14) magnesium (s) + oxygen (g) \text{heat} → magnesium oxide (s)

15) nitric acid (aq) +calcium hydroxide (s) → calcium nitrate (aq) + water

16) sodium hydroxide (aq) + copper(II) nitrate (aq) → sodium nitrate (aq) + copper (II) hydroxide (s)

17) strontium iodide (aq) + sodium phosphate (aq) → strontium phosphate (s) + sodium iodide (aq)

18) lithium carbonate (aq) + hydroiodic acid (aq) → lithium iodide (aq) + carbon dioxide + H_2O(l)

19) potassium oxalate (aq) + silver acetate (aq) → potassium acetate (aq) + silver oxalate (s)

20) iron (III) sulfate (aq) + calcium chlorate (aq) → iron (III) chlorate (aq) + calcium sulfate (s)

21) ammonium bicarbonate (s) + heat → ammonia (g) + carbon dioxide (g) + water (g)
Worksheet 5: Stoichiometry 1

1. How many moles of Nitrogen atoms are in 1 mole of N₂O₅? Name this compound.
2. How many moles of Nitrogen atoms are in 2 mole of N₂O₅?
3. How many moles of Nitrogen atoms are in 8.6 mol of N₂O₅?
4. How many moles of Nitrogen atoms are in 8.6 g N₂O₅?
5. How many Pb atoms are in 1.00×10⁻⁹ g of Pb metal?
6. How many g are in 1 mol of Na₂CO₃ (name this compound also)?
7. How many moles of H₂SO₄ are present in 45.8 g of the substance? (name?)
8. What is the ratio of H atoms to C atoms in 1 molecule of C₂H₆?
9. If one has 44.3 g of C₂H₆, how many moles of C and of H atoms are present?
10. Mercury is a liquid at room temperature, and has a density of 13.5 g/mL. How many mol of Hg are in 2.75 mL of Hg?
11. Like many metals, manganese reacts with fluorine gas to make a salt:
   \[ \text{Mn(s)} + \text{F}_2(g) \rightarrow \text{MnF}_3(s) \]
   a. If 3.70 mol Mn is reacted with excess F₂, how many mol of MnF₃ will form?
   b. If 1.25×10⁻³ mol of MnF₃ form, how many mol of F₂ was consumed?
   c. How many g of MnF₃ can be made from 1.67 g of F₂?
12. Write balanced chemical equations for the following:
   a. In a gaseous reaction, dihydrogen sulfide burns in oxygen to form sulfur dioxide and water vapor.
   b. Hydrogen gas is passed over powdered iron (II) oxide. Iron metal and water vapor form.
   c. Iron (II) chloride can be converted to iron (III) fluoride by reacting it with chlorine trifluoride gas. Chlorine gas is also produced in this reaction.
13. Calculate each of the following from the equations written in #12 above.
   a. How many grams of oxygen will react with 3.22 g hydrogen sulfide?
   b. How many grams of iron metal are formed if 16.754 g of iron (II) oxide reacts with excess hydrogen gas?
   c. How many grams of chlorine gas will be produced if 27 g of iron (II) chloride reacts with excess chlorine trifluoride?
Worksheet 6: Stoichiometry 2

1. If you had 1 mole of $100 bills, and you wanted to spend them at the leisurely rate of $50,000,000 per second, how many years would it take to spend the money?

2. In the reaction with Al and H₂SO₄ below, what mass of aluminum is needed to ensure that 62.7 g H₂SO₄ is completely used up.

   \[ \text{Al(s)} + \text{H}_2\text{SO}_4\ (aq) \rightarrow \text{Al}_2(\text{SO}_4)_3\ (aq) + \text{H}_2\ (g) \]  

   [unbalanced]

3. Before starting this problem, take a guess at the amount of carbon dioxide gas that is produced by your car for each gallon of gasoline that you consume. For comparison, water has a mass of about 8 pounds per gallon, and gasoline is about 80% as heavy. Now, calculate the mass of CO₂ (g) that is produced from burning 1 gallon of gasoline. The density of gasoline is 0.80 g/mL. A generic formula for gasoline is C₈H₁₈.

4. Nitrogen monoxide can be prepared by the oxidation of ammonia by the following equation:

   \[ \text{NH}_3\ (g) + \text{O}_2\ (g) \rightarrow \text{NO}\ (g) + \text{H}_2\text{O}\ (g) \]

   a. If 45.7 g of NH₃ and 52.5 g of O₂ react together, how many g of NO will be formed?

   b. How many grams of water you would expect to be formed in a reaction with 22.3 g of oxygen and 16.2 g of ammonia.

   c. If 42.3 mg of NO is formed from NH₃ and O₂, how many g of each reactant was used?

5. How much product is made when 42.5 g of Mg metal reacts with 56.9 g of oxygen gas? How much of the reactant in excess is left over?

6. How much (in g) of AgCl (s) can be made by mixing 20.0 mL of AgNO₃ (2.323 g/L) with 50.0 mL of KCl (0.982 g/L)? How many ions (in mol) of the excess reactant remains after the reaction?

7. Elemental phosphorus (a solid) reacts with elemental chlorine to make phosphorus trichloride. How many g of each reactant is needed to make 100.0 g of phosphorus trichloride (a liquid).

8. Potassium chlorate decomposes when heated to make potassium chloride and oxygen gas. How many g of the reactant is needed to make 16.0 g of molecular oxygen?

9. Aspirin has the chemical name acetylsalicylic acid. It has the following composition: C 60.01%, H 4.48%, O 35.52%. Determine the empirical formula for aspirin.

10. Aspirin has a formula weight of 180.16. What is the molecular formula for aspirin?

11. Serotonin (FW 176 g/mol) is a compound that conducts nerve impulses in brain and muscle tissue. It contains 68.2 mass % carbon, 6.86% H, 15.9 % N and 9.08% O. What is the molecular formula for serotonin?
Worksheet 7: Concentration 1

1. What is the molar concentration of a solution prepared by diluting 3.18 g sodium nitrate to a final volume of 150.0 mL?

2. How many grams of potassium sulfate would be required to make 235.0 mL of 0.152 M solution?

3. What volume of 0.320 M iron (III) chloride can be prepared with 21.7 g of the solid?

4. How much (in L) 0.200 M solution can be prepared from 150.0 mL of 1.69 M KBr?

5. What is the molarity of a solution if 3.79 g of cobalt (II) nitrate hexahydrate is dissolved in water to make 100.0 mL of solution?

6. A solution is made by dissolving 12.98 g of ammonium phosphate in water to a final volume of 225.0 mL. What is the molar concentration of the resulting solution?

7. If 15.0 mL of the solution from the previous problem is diluted to 75.0 mL, what is the concentration of the diluted solution?

8. An experiment calls for 275 mL of 2.00 M HCl. The only solution HCl available is 5.89 M. Write the steps and calculations needed to prepare the 2.00 M solution.

9. What is the molarity of a concentrated sulfuric acid solution if 35.0 mL of this solution is diluted to a final volume of 200.0 mL, producing a 0.145 M solution?

10. How many mL of concentrated nitric acid (16.0 M) would you need to dilute to 300.0 mL to make a 0.550 M solution?

11. How many mL of 3.00 M HCl would need to be diluted in order to make 175.0 mL of 0.125 M HCl?

12. How would you prepare 250.0 mL of 0.0973 M NH₄Br solution, starting with a 0.230 M solution? How many moles of NH₄Br are present in 100 mL of the dilute solution? What is the new concentration if 50.0 mL of the 0.0973 M solution was diluted with 44.0 mL H₂O?

13. Your instructor asks you to prepare 155 mL of a saline solution that is 9.0 g of NaCl per liter of water. The only source of NaCl that you can find in the lab is 62 mL of 0.500 M solution. Is this enough to make up the desired saline solution? How much extra volume do you have, or how much extra volume do you need?

14. 100.00 mL CuSO₄ (aq) is diluted to 500 mL with water. The diluted portion is found to have 5.11 g/L of Cu²⁺ ions. What is the molar concentration of CuSO₄ in the original 100 mL?

15. What is the molarity of water? The density of water is 1.00 g/mL
Worksheet 8: Practice Problems from Chapter 4 and 16

1. Calculate the molarity of 10.00 mL HCl titrated with 17.05 mL of 0.1245 M NaOH

2. What is the molarity of a phosphoric acid solution if 12.00 mL of the solution is neutralized by 23.97 mL of 0.2500 M sodium hydroxide?

3. If aqueous potassium sulfide is mixed with aqueous aluminum acetate, will a reaction occur? If so, write the molecular, total ionic, and net ionic equations.

4. What is the concentration of acetate ions in 0.300 M aluminum acetate? What is the concentration of sulfide ions in 0.170 M potassium sulfide?

5. What is the maximum amount of solid product that will form when 50.0 mL of 0.300 M aluminum acetate is mixed with 100 mL of 0.170 M potassium sulfide? Which ions remain in solution?

6. How many mL of a 0.250 M magnesium chloride solution are needed to react completely with 35.0 mL of 0.100 M lead (II) nitrate?

7. What volume of 0.249 M H₂SO₄ is needed to react with 23.97 mL of 1.04 M NaOH?

8. 14.6 g of calcium chloride is dissolved in water to make 1.00 L of solution. 65.6 g of silver nitrate is dissolved in water to make 1.50 L of solution. If 25.0 mL of each of these two solutions are mixed together, what is the maximum amount of precipitate that will form?

9. How many mL of 3.00 M HCl would need to be diluted in order to make 175.0 mL of 0.125 M HCl?

10. Which solution is more acidic, one with pH = 3 or pH = 5? Which one has a higher [H⁺]?

11. What is the pH of a 0.0079 M solution of HNO₃. What is the pOH of this solution?

12. If a solution has pH = 3.26, what is the [H₃O⁺] in moles/L?

13. Draw a picture showing why 0.12 M HCl has a lower pH than 0.12 M H₂C₂O₄.

14. What is the pH of the resulting solution if 0.0376 L of 0.12 M HCl is mixed with 0.0619 L of 0.071 M NaOH?

15. Oxalic acid (H₂C₂O₄) is a toxic compound found in rhubarb leaves. To find out how much oxalic acid is present in a sample of rhubarb, a student takes 195 g of leaves and brews them (like tea) in 500.0 mL of H₂O. Using a pipet, 25.00 mL of this solution is removed and placed in a small flask. The acid is then titrated with 18.0 mL of 0.109 M NaOH. Determine the percentage by mass of oxalic acid in the leaves.
Worksheet 9: Thermodynamics 1

1. Fill in the states of matter in each box, and the terms sublimation, vaporization, condensation, melting (fusion), freezing, deposition.

2. Determine if energy is absorbed or released, and whether each process is endo or exothermic.
   - Melting ice requires ice to ____________________ energy; ____________________
   - Boiling water requires water to ________________ energy;  ____________________
   - Condensing steam to water _________________ energy;  ____________________
   - Freezing water to ice ______________________ energy; ____________________

3. Calculate the specific heat of each substance using the information given. Do not look up the values in your book. Does it take the same amount of heat to raise 1 gram of the substance by 1 °C?

<table>
<thead>
<tr>
<th>Substance</th>
<th>Heat added</th>
<th>Mass</th>
<th>Temp change</th>
<th>Specific Heat J/g °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>water</td>
<td>55.5 J</td>
<td>2.09 g</td>
<td>6.35 °C</td>
<td></td>
</tr>
<tr>
<td>Iron</td>
<td>3.01 kJ</td>
<td>38.9 g</td>
<td>172 °C</td>
<td></td>
</tr>
<tr>
<td>Mercury</td>
<td>451 J</td>
<td>111.1 g</td>
<td>29.0 °C</td>
<td></td>
</tr>
<tr>
<td>NaCl (s)</td>
<td>132 J</td>
<td>8.89 g</td>
<td>106 °C</td>
<td></td>
</tr>
</tbody>
</table>

4. How much heat is required to raise 335 mL of water from 20.0 °C to 95.5 °C?

5. A 55.1 g piece of metal is heated to a temperature of 45.1 °C, and placed into a cup containing 359 g of water at 20.0 °C. The final temperature of the water and metal is 22.3°C.
   a. How much heat energy did the water absorb?
   b. How much heat energy did the metal release to the water?
   c. What is the specific heat of the metal?

6. There is an attractive force between a rock and the earth (gravity). Energy is released when:
   a. the rock is lifted above the ground
   b. the rock falls and hits the ground
   c. whenever the rock is moved

7. There is an attractive force between a magnet and your fridge door. Energy is released when:
a. the magnet is pulled away from the fridge
b. the magnet is pulled out of your hand toward the fridge
c. whenever the magnet moves

8. There is an attractive force between atoms in a compound. Chemical energy is released when:
   a. chemical bonds are broken
   b. chemical bonds are formed
   c. during any chemical reaction

9. Is energy released or absorbed by the following processes?
   a. H–H bond breaks, separating two atoms with an attractive force, __________________
   b. F–F bond breaks, separating two atoms with an attractive force, __________________
   c. 2 H–F bonds form, joining two atoms with an attractive force, ________________
Worksheet 10: Thermodynamics 2

1. Given the reaction of \( \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (l) \) : \( \Delta H^\circ = -890 \text{ kJ} \)
   a. is the reaction endo or exothermic?
   b. does the reaction release heat to or absorb heat from the surroundings?
   c. how much heat energy is released if 1 mole of \( \text{CO}_2 \) is formed?
   d. how much heat energy is released if 2 mole of \( \text{CO}_2 \) is formed?
   e. how much heat energy is released if 0.054 mole of \( \text{CO}_2 \) is formed?

2. Given the reaction equation of \( 2 \text{SO}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{SO}_3 (g) \) \( \Delta H^\circ = -192.8 \text{ kJ} \):
   1. Calculate \( \Delta H^\circ \) for a reaction that produces \( 2.00 \times 10^2 \text{ g of SO}_3 \).
   2. Is the reaction endothermic or exothermic? What does this mean?

3. A 36.9 g sample of metal is heated to 100.0 °C, and then added to a calorimeter containing 141.5 g of water at 23.1 °C. The temp. of the water rises to a max of 25.2 °C before cooling back down.
   a. Did the water absorb heat or did it release heat?
   b. How many joules of heat was exchanged between the water and the metal?
   c. What is the identity of the metal?

4. Methanol is a clean burning liquid fuel used in dragsters and it may one day be the fuel that is used for all cars. It can be made from methane found in natural gas:
   \[ 2 \text{CH}_4 (g) + \text{O}_2 (g) \rightarrow 2 \text{CH}_3\text{OH} (l) \]
   Find the \( \Delta H^\circ \) for the formation of methanol using the following equations:
   \[ \text{CH}_4 (g) + \text{H}_2\text{O} (g) \rightarrow 3 \text{H}_2 (g) + \text{CO} (g) \] \( \Delta H^\circ = + 206 \text{ kJ} \)
   \[ 2 \text{H}_2 (g) + \text{CO} (g) \rightarrow \text{CH}_3\text{OH} (l) \] \( \Delta H^\circ = - 128 \text{ kJ} \)
   \[ 2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O} (g) \] \( \Delta H^\circ = - 484 \text{ kJ} \)

5. Calculate the amount of energy gained in burning 1 mole of ethanol, \( \text{C}_2\text{H}_5\text{OH} \) in \( \text{O}_2 \), and for 1 mole of methanol \( \text{CH}_3\text{OH} \) in \( \text{O}_2 \) using \( \Delta H_f^\circ \) values from your book.

6. Use the information in the previous question to compare the energy density (kJ/g) of each substance. Describe which substance would you choose as a fuel for your car.

7. Write the equation that represents the \( \Delta H_f^\circ \) for: a. \( \text{SO}_2(s) \)  b. \( \text{HCl}(g) \)  c. \( \text{C}_2\text{H}_5\text{OH}(l) \)

8. Iron metal is produced through a series of complex reactions. The overall reaction is:
   \[ \text{Fe}_2\text{O}_3(s) + 3 \text{CO}(g) \rightarrow 2 \text{Fe}(s) + 3 \text{CO}_2(g) \]
   Find the \( \Delta H^\circ \) for the formation of iron using the following equations (do not use your book!):
   \[ 3 \text{Fe}_2\text{O}_3 (s) + \text{CO}(g) \rightarrow 2 \text{Fe}_3\text{O}_4(s) + \text{CO}_2(g) \] \( \Delta H^\circ = -48.5 \text{ kJ} \)
   \[ \text{Fe}(s) + \text{CO}_2(g) \rightarrow \text{FeO(s)} + \text{CO}(g) \] \( \Delta H^\circ = -11.0 \text{ kJ} \)
   \[ \text{Fe}_3\text{O}_4 (s) + \text{CO}(g) \rightarrow 3 \text{FeO(s)} + \text{CO}_2(g) \] \( \Delta H^\circ = + 22.0 \text{ kJ} \)
Worksheet 11: Thermodynamics 3

1. The standard heat of formation (\(\Delta H_f\)) for carbon monoxide gas is \(-110.5\) kJ/mol. Write the thermochemical equation for the formation of carbon monoxide.

2. The heat capacity for a bomb calorimeter is \(5.15\) kJ/°C. When \(0.480\) g \(H_2(g)\) is burned in the calorimeter, the temperature of the calorimeter increased from \(25.00\) °C to \(36.16\) °C. Calculate the \(\Delta H^\circ\) for this reaction.

\[
H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)
\]

3. Use \(\Delta H_f\) values to calculate the \(\Delta H^\circ\) for the following reaction. \(CO_2(g); \Delta H_f = -393.5\) kJ/mol (see above for CO).

\[
2 \ CO(g) + O_2(g) \rightarrow 2 \ CO_2(g)
\]

4. Use the \(\Delta H^\circ\) calculated for the three reactions above and the additional \(\Delta H^\circ\) provided below to calculate the \(\Delta H^\circ\) for the reaction between \(CO\) gas and hydrogen gas, producing methane gas (\(CH_4\)), \(CO_2\) gas and water vapor. Balance the equation where 3 moles of \(CO\) gas reacts.

\[
C(\text{graphite}) + 2 \ H_2(g) \rightarrow CH_4(g) \quad \Delta H^\circ = -74.81\ \text{kJ}
\]

5. A student heats \(21.90\) g of \(Cu\) metal to \(90.0\) °C. This metal is then placed into \(98.26\) g of water at \(25.7\) °C in a calorimeter. The mixture is stirred until the temperature of the water remains unchanged at \(27.0\) °C.

What is the final temperature of the copper?

How much heat energy does the copper release, based on data specific to copper?

How much heat energy does the water absorb, based on data specific to water?

6. The student repeats the experiment with \(42.5\) g of an unknown metal. The metal is heated again to \(90.0\) °C, and added to \(102.4\)g of \(25.6\) °C water in the calorimeter. The final temperature becomes \(27.4\) °C. What is the specific heat of the metal?
Worksheet 12: Chapter 6

What are the symbols for:

frequency: ________, wavelength __________

definition of ν is # of ______________________________________________________

Units of ν is __________________________________________

• speed of wave, c, = frequency * wavelength formula: __________________

• For light in vacuum, speed = c = _________________

• Einstein: light travels in energy packets, called __________________

• A photon is a ____________________, in terms of E

• Formula for the E of one photon: __________________

Colors are emitted from excited atoms when __________________________________________

• Bohr: _________________ travel in _________________

• Light is emitted when e– move _____________________________________________

• E is ____________________________, __________________ emitted from excited atoms ___________________________

• emitted light appears_____________________________________ (single energies)

Lecture questions:

1. What is the wavelength of light that has a ν of 7.95×10^{14} s\(^{-1}\)?

2. What is the ν of light with λ = 3.33 m?

3. What is the energy of light that has a ν of 7.95×10^{14} s\(^{-1}\)?

4. What is the energy of a photon with λ = 3.33 m?

5. How much energy (in kJ) is present in 4.56×10^{-2} moles of photons with λ = 551 nm?
Worksheet 13: Chapter 7

1. Calculate $Z_{eff}$ for:

<table>
<thead>
<tr>
<th></th>
<th>$Z-s = Z_{eff}$</th>
<th>Mg</th>
<th>Cl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>11-10 = 1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>K</td>
<td></td>
<td>Ca</td>
<td>Br</td>
</tr>
<tr>
<td>Rb</td>
<td></td>
<td>Sr</td>
<td>I</td>
</tr>
</tbody>
</table>

$Z =$ atomic number    $s =$ shielding or core electrons

2. Which of the following atoms is larger: Mg Na Cl?

3. Which is bigger in this isoelectronic series?

$O^{2-}$  $F^{-}$  $Na^{+}$  $Mg^{2+}$  $Al^{3+}$

4. According to $e^-$ configurations, from which orbitals are the following $e^-$ lost?

<table>
<thead>
<tr>
<th>Element</th>
<th>1st $e^-$</th>
<th>2nd $e^-$</th>
<th>3rd $e^-$</th>
<th>4th $e^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td></td>
<td></td>
<td>X</td>
<td>X</td>
</tr>
<tr>
<td>Be</td>
<td></td>
<td></td>
<td></td>
<td>X</td>
</tr>
<tr>
<td>B</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Electron configuration diagram for multi-electron atoms
D orbital sets drop in Energy upon filling
Worksheet 14: VSEPR and Valence Bond Theory

1. For each of the molecules shown below, give the name of the shape, the hybridization on the central atom and the bond angles. Assuming that the bonds are polar, are the molecules polar?

   - shape _____________ _____________ _____________ _____________
   - hybridization _____________ _____________ _____________ _____________
   - bond angles _____________ _____________ _____________ _____________
   - polar? _____________ _____________ _____________ _____________

2. In the molecule shown below, for each atom labeled with a number, predict the shape, hybridization and angle. Also predict the types of bonds (σ or π) listed in the last column.

   ![Molecule](image)

<table>
<thead>
<tr>
<th>atom number</th>
<th>shape</th>
<th>hybridization</th>
<th>angle</th>
<th>bond types</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
<td>H–N₁</td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td>N₁–C₂</td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td>C₂–C₃</td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
<td>C₃–C₄</td>
</tr>
<tr>
<td>5</td>
<td></td>
<td></td>
<td></td>
<td>C₄–C₅</td>
</tr>
</tbody>
</table>

3. For the following bonds, describe what type of orbitals on each atom overlap to form any σ or π bonds, according to VBT. double bond in the molecule above

   N₁=C₂  C₃–C₄  C₄–C₅
Worksheet 15: Molecular Orbital Theory  

In Molecular orbital theory:
Atomic orbitals overlap to create molecular orbitals
For a diatomic species (N₂, O₂), the atomic orbitals from each atom are represented in the MO diagram, and each atomic orbital that mixes with another similar atomic orbital creates a pair of MO’s.

• For every _______________ AO’s, 2 _______________ MO’s form
  • 1 higher in E
  • 1 lower in E

• all e⁻ from both atoms go into the MO’s
• Bonding e⁻ are in MO’s _______________ in E than AO’s
• Anti bonding e⁻ are in MO’s _______________ in E than AO’s

Label, for H₂ and He₂:
1. boxes for atomic orbitals
2. box for bonding orbital
3. box for antibonding orbital
4. Relative energies

For Li₂ and Be₂:
1. The bond length in $\text{C}_2$ is measured by neutron diffraction to be 131 pm (100 pm = 1Å). Using your information about carbon-carbon bonds from the computer modeling lab, does the $\text{C}_2$ bond resemble a $\text{C}–\text{C}$, a $\text{C} = \text{C}$, or a $\text{C}≡\text{C}$?

2. Predict the bond order of $\text{C}_2$ via MO theory.

3. Calcium and carbon form an ionic compound, $\text{CaC}_2$. Draw the lewis structure of the $\text{C}_2$ anion. What are the MO orbital configuration and the bond order of this anion? Do VBT and MO theory agree?

4. The ionization energy of $\text{O}_2$ is lower than the IE of atomic $\text{O}$. Explain this in terms of the where the electrons are in atomic and MO’s of $\text{O}_2$.

5. Do you expect the IE of $\text{N}_2$ to be higher or lower than that of atomic $\text{N}$?

6. What is the effect on the bond order of chemically reducing $\text{N}_2$ with one electron? 2 electrons?

7. What is the effect on the magnetic properties of $\text{N}_2$ when it is chemically reduced with one electron? Two electrons? Three electrons?

For $\text{B, C, and N}$

For $\text{O, F and Ne}$
Worksheet 17: Gases

1. A sample of gas in a sealed steel container is heated from 25 °C to 50 °C. The gas was initially at a pressure of 2.35 atm. What is the new pressure?

2. If a sample of gas at 1.00 atm and 308 K is allowed to expand at constant temperature from 199 mL to 398 mL, what is the new pressure of the gas?

3. How many moles are in a sample of N₂ gas in a container that holds 35 L, with a pressure of 1.1 atm, and a temperature of 298 K?

4. CaCO₃(s) decomposes to calcium oxide(s) and CO₂ (g). If a sample of CaCO₃ is decomposed, and the CO₂ is collected in a 250 mL flask at 1.3 atm pressure and 31 °C, how many g of CaCO₃ were used?

5. NaN₃ (sodium azide) is a solid used in air bags. It decomposes to Na(s) and N₂ (g). If a Lexus air bag holds 45.5 L of gas, and the temperature is 22.0 °C at a pressure of 1.09 atm, how many g of NaN₃ is needed to charge an airbag?

6. How many mL of gas are formed when 15.5 mL of 1.44 M sulfuric acid is mixed with 25.7 mL of 0.77 M sodium bicarbonate? The reaction occurs at 20.0 °C and 0.980 atm pressure.

7. Lithium metal reacts with heavy water (D₂O, where D is the symbol for Deuterium) to make D₂ (g). If 8.125 g of Li reacts with 16.7 g of D₂O, what is the pressure of the D₂ gas formed if it occupies 1.45 L at 295 K?

8. If 10.0 g of O₂ is mixed with 5.00 g of CH₄ in a 10 L container at 25 °C, what is the partial pressure of each gas, and the total pressure of the container?

9. While cleaning his bathroom, a distracted student mixes 650 mL of bleach (5.1% wt/vol sodium hypochlorite) with a full bottle (1 pt) of hydrogen peroxide (3.1% wt/vol). A violent reaction occurs, forming a large amount of a colorless gas. How many L of what gas are formed in this reaction?

10. Water and CCl₄ (a liquid, is it polar or non-polar?) do not mix. Vapors from water and CCl₄ mix just fine however. Can you explain this?

11. Use circles to represent H₂O molecules and draw pictures to show how the molecules would be arranged under the conditions listed for each box.

   ![Gas-Phase Water](image1)
   ![Liquid Water](image2)
   ![Solid Water (Ice)](image3)

12. How does a gas differ from a liquid in the following properties:
   a. number of particles per volume   b. Compressibility   c. density

13. Why do hot air balloons float?