Summer Reading 2015

Upper School Science

Advanced Placement Chemistry

Ms. Scanlon

AP Chemistry is a difficult course and the AP Exam is very challenging. To succeed, you must keep up with the assignments and be willing to spend time working through the material. Like most AP classes, AP Chemistry comes with a summer assignment. I encourage you to form a study group and begin by working on the summer assignment together (several local stores and restaurants have WiFi…). The course is much easier if you have a support system. If you have problems, please feel free to e-mail me at christine.scanlon@woosterschool.org. (I will check my e-mail infrequently at times during the summer, but I will get back to you!) Please take the summer assignment seriously.

Part 1 is a couple of AP Chemistry questions that use material that was covered in your first year chemistry class. The purpose of these questions is for you to review this material and to come into class in September ready to start on new material. The AP Chemistry exam both includes all the material covered in the first year chemistry class and directly builds on that material, so a solid understanding of the these topics is critical.

Part 2 is a list of information that must be memorized. All of the material I am giving you must be memorized before you take the AP exam, as it is not provided for you. Getting the route memorization out of the way will free up time during the year for us to focus on the new material more deeply and to complete more labs. In addition, being able to recall this information will make many of the more complex new topics easier to address.

Part 3 is an optional reading assignment. The book, *The Tale of Seven Elements*, by Scerri, is a great read for many reasons. It provides context to the chemistry principles we will study. It reviews many important topics such as the organization of the periodic table and atomic structure. It is a great story that provides insight into the great complexity of scientific research and attribution that is usually completely omitted from textbooks.

Have a great summer,

Ms. Scanlon

Part 1: AP Chemistry Review Questions

These questions are to be completed and ready to discuss on the first day of class. I will collect them, and they will be graded for completeness, not correctness. Use your old chemistry textbook, new AP Chemistry textbook, and online review sources such as Crash Course Chemistry and Socratic.org/chemistry. Write down your work for each problem. Use additional paper – do not try to squeeze your writing in between these questions. Do NOT search for the answer keys online.

Question 1: Stoichiometry and EM radiation
Question 2: Isotopes and Atomic Mass
Question 3: EM radiation and the Atom
Question 4: Periodic table, Elements and Compounds, Electron Configurations
Question 5: Periodic Trends and Electron Configurations
Question 6: Periodic Trends
Question 7: Stoichiometry and the Ideal Gas Law
Question 8: Gas Laws, Density, and Moles
Question 9: Lewis Structures, VSEPR, and Polarity

$$CH_4(g) + 2\ Cl_2(g) \rightarrow CH_2Cl_2(g) + 2\ HCl(g)$$

3. Methane gas reacts with chlorine gas to form dichloromethane and hydrogen chloride, as represented by the equation above.

(a) A 25.0 g sample of methane gas is placed in a reaction vessel containing 2.58 mol of $Cl_2(g)$.

(i) Identify the limiting reactant when the methane and chlorine gases are combined. Justify your answer with a calculation.

(ii) Calculate the total number of moles of $CH_2Cl_2(g)$ in the container after the limiting reactant has been totally consumed.

Initiating most reactions involving chlorine gas involves breaking the Cl–Cl bond, which has a bond energy of 242 kJ mol^{-1}.

(b) Calculate the amount of energy, in joules, needed to break a single Cl–Cl bond.

(c) Calculate the longest wavelength of light, in meters, that can supply the energy per photon necessary to break the Cl–Cl bond.

2. Answer the following problems about gases.

(a) The average atomic mass of naturally occurring neon is 20.18 amu. There are two common isotopes of naturally occurring neon as indicated in the table below.

Isotope	Mass (amu)
Ne-20	19.99
Ne-22	21.99

(i) Using the information above, calculate the percent abundance of each isotope.

(ii) Calculate the number of Ne-22 atoms in a 12.55 g sample of naturally occurring neon.

(b) A major line in the emission spectrum of neon corresponds to a frequency of $4.34 \times 10^{14}\ s^{-1}$. Calculate the wavelength, in nanometers, of light that corresponds to this line.

(c) In the upper atmosphere, ozone molecules decompose as they absorb ultraviolet (UV) radiation, as shown by the equation below. Ozone serves to block harmful ultraviolet radiation that comes from the Sun.

$$O_3(g) \xrightarrow{UV} O_2(g) + O(g)$$

A molecule of $O_3(g)$ absorbs a photon with a frequency of $1.00 \times 10^{15}\ s^{-1}$.

(i) How much energy, in joules, does the $O_3(g)$ molecule absorb per photon?

(ii) The minimum energy needed to break an oxygen-oxygen bond in ozone is 387 kJ mol^{-1}. Does a photon with a frequency of $1.00 \times 10^{15}\ s^{-1}$ have enough energy to break this bond? Support your answer with a calculation.

2. Answer the following questions regarding light and its interactions with molecules, atoms, and ions.

(a) The longest wavelength of light with enough energy to break the Cl–Cl bond in $Cl_2(g)$ is 495 nm.

(i) Calculate the frequency, in s^{-1}, of the light.

(ii) Calculate the energy, in J, of a photon of the light.

(iii) Calculate the minimum energy, in kJ mol^{-1}, of the Cl–Cl bond.

(b) A certain line in the spectrum of atomic hydrogen is associated with the electronic transition in the H atom from the sixth energy level ($n = 6$) to the second energy level ($n = 2$).

(i) Indicate whether the H atom emits energy or whether it absorbs energy during the transition. Justify your answer.

(ii) Calculate the wavelength, in nm, of the radiation associated with the spectral line.

(iii) Account for the observation that the amount of energy associated with the same electronic transition ($n = 6$ to $n = 2$) in the He^{+} ion is greater than that associated with the corresponding transition in the H atom.

8. Suppose that a stable element with atomic number 119, symbol Q, has been discovered.

(a) Write the ground-state electron configuration for Q, showing only the valence-shell electrons.

(b) Would Q be a metal or a nonmetal? Explain in terms of electron configuration.

(c) On the basis of periodic trends, would Q have the largest atomic radius in its group or would it have the smallest? Explain in terms of electronic structure.

(d) What would be the most likely charge of the Q ion in stable ionic compounds?

(e) Write a balanced equation that would represent the reaction of Q with water.

(f) Assume that Q reacts to form a carbonate compound.

(i) Write the formula for the compound formed between Q and the carbonate ion, CO_3^{2-}.

(ii) Predict whether or not the compound would be soluble in water. Explain your reasoning.

	First Ionization Energy ($kJ\ mol^{-1}$)	Second Ionization Energy ($kJ\ mol^{-1}$)	Third Ionization Energy ($kJ\ mol^{-1}$)
Element 1	1,251	2,300	3,820
Element 2	496	4,560	6,910
Element 3	738	1,450	7,730
Element 4	1,000	2,250	3,360

6. The table above shows the first three ionization energies for atoms of four elements from the third period of the periodic table. The elements are numbered randomly. Use the information in the table to answer the following questions.

(a) Which element is most metallic in character? Explain your reasoning.

(b) Identify element 3. Explain your reasoning.

(c) Write the complete electron configuration for an atom of element 3.

(d) What is the expected oxidation state for the most common ion of element 2 ?

(e) What is the chemical symbol for element 2 ?

(f) A neutral atom of which of the four elements has the smallest radius?

Answer Question 5 and Question 6. The Section II score weighting for these questions is 15 percent each.

Your responses to these questions will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

5. Using principles of atomic and molecular structure and the information in the table below, answer the following questions about atomic fluorine, oxygen, and xenon, as well as some of their compounds.

Atom	First Ionization Energy ($kJ\ mol^{-1}$)
F	1,681.0
O	1,313.9
Xe	?

(a) Write the equation for the ionization of atomic fluorine that requires 1,681.0 $kJ\ mol^{-1}$.

(b) Account for the fact that the first ionization energy of atomic fluorine is greater than that of atomic oxygen. (You must discuss both atoms in your response.)

(c) Predict whether the first ionization energy of atomic xenon is greater than, less than, or equal to the first ionization energy of atomic fluorine. Justify your prediction.

3. Answer the following questions that relate to the analysis of chemical compounds.

(a) A compound containing the elements C, H, N, and O is analyzed. When a 1.2359 g sample is burned in excess oxygen, 2.241 g of $CO_2(g)$ is formed. The combustion analysis also showed that the sample contained 0.0648 g of H.

(i) Determine the mass, in grams, of C in the 1.2359 g sample of the compound.

(ii) When the compound is analyzed for N content only, the mass percent of N is found to be 28.84 percent. Determine the mass, in grams, of N in the original 1.2359 g sample of the compound.

(iii) Determine the mass, in grams, of O in the original 1.2359 g sample of the compound.

(iv) Determine the empirical formula of the compound.

(b) A different compound, which has the empirical formula CH_2Br, has a vapor density of 6.00 $g\ L^{-1}$ at 375 K and 0.983 atm. Using these data, determine the following.

(i) The molar mass of the compound

(ii) The molecular formula of the compound

2. A rigid 8.20 L flask contains a mixture of 2.50 moles of H_2, 0.500 mole of O_2, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127°C.

   (a) Calculate the total pressure in the flask.

   (b) Calculate the mole fraction of H_2 in the flask.

   (c) Calculate the density (in g L^{-1}) of the mixture in the flask.

   The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

$$2\,H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

   (d) Give the mole fraction of all species present in the flask at the end of the reaction.

8. Answer the following questions using principles of chemical bonding and molecular structure.

   (a) Consider the carbon dioxide molecule, CO_2, and the carbonate ion, CO_3^{2-}.

      (i) Draw the complete Lewis electron-dot structure for each species.

      (ii) Account for the fact that the carbon-oxygen bond length in CO_3^{2-} is greater than the carbon-oxygen bond length in CO_2.

   (b) Consider the molecules CF_4 and SF_4.

      (i) Draw the complete Lewis electron-dot structure for each molecule.

      (ii) In terms of molecular geometry, account for the fact that the CF_4 molecule is nonpolar, whereas the SF_4 molecule is polar.

Part 2: Items to Memorize over the Summer

You will have a quiz on the first week of class on the following information. You MUST memorize these. Get an early start! Try using flashcards, or make up mnemonic devices! It may help you to write the information over again and again, or to read the information to yourself again and again. Try studying "actively." Learn about yourself and discover how you remember best. Learning to study more effectively now will help you in college later! Here is a website with many games that aid in memorizing some of these topics: http://www2.stetson.edu/mahjongchem/.

1. Common Ions and Polyatomic Ions: Memorize the list. I have also attached flash cards with the polyatomic ions that you can print double sided and then cut apart.

2. Solubility Rules: Memorize the list.

3. Acids and Bases: Memorize the list of common acids and bases as well as the list of strong acids and bases. Know both the name and the formulas for the chemicals.

4. Names and Symbols of the elements: The periodic table provided for you to use on the AP exam does not have the names of the elements, but just the symbols. Study so that if you are given the element symbol, you automatically know the name, and if you are given the element name, you automatically know the symbol. (You should know most of them already ☺)

5. Diatomic elements: ClOBrFINH (Chlorine, Oxygen, Bromine, Fluorine, Iodine, Nitrogen, Hydrogen). Recalling that these (except for hydrogen) make an upside down L on the right side of the periodic table can help you here.

6. Significant Figures Rules: Know the rules for counting significant figures and the rules for calculating with significant figures.

7. Procedures: This is a list of little skills needed to solve problems. Usually these are things you do before the first step in a problem. These should be routine.
    a. Calculate the molar mass of an element of compound
    b. Convert between units, specifically
        i. Metric system
        ii. Grams, moles, Avogadro's number, and volume of gas at STP
    c. Balance chemical equations

8. Colors:

Flame Test Colors:
Barium – green
Calcium – orange
Copper – blue/green
Lithium – red
Potassium – lavender
Sodium – yellow/orange
Strontium – orange/red

Colors of Some Aqueous Ions

Co^{2+}	pink
Cu^{2+}	blue green
Fe^{3+}	olive green
Ni^{2+}	bright green
Fe^{3+}	brown
CrO_4^{2-}	orange
$Cr_2O_7^{2-}$	yellow

1. Common Ions

Positive Ions (cations)	**Negative ions (anions)**
+1 Charge	***-1 Charge***
Ammonium (NH_4^+)	Acetate ($C_2H_3O_2^-$)
Copper (I) or cuprous (Cu^+)	Cyanide (CN^-)

Hydrogen (H^+) "proton"	Dihydrogen phosphate ($H_2PO_4^-$)
Hydronium ion (H_3O^+)	Hydrogen carbonate or bicarbonate (HCO^{3-})
Silver (Ag^+)	Hydrogen sulfate or bisulfate (HSO_4^-)
Group 1 (Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+)	Hydroxide (OH^-)
	Nitrate (NO_3^-)
+2 Charge	Nitrite (NO_2^-)
Cadmium (Cd^{2+})	Perchlorate (ClO_4^-)
Chromium(II) or chromous (Cr^{2+})	Chlorate (ClO_3^-)
Cobalt(II) or cobaltous (Co^{2+})	Chlorite (ClO_2^-)
Copper(II) or cupric (Cu^{2+})	Hypochlorite (ClO^-)
Iron(II) or ferrous (Fe^{2+})	Permanganate (MnO_4^-)
Lead(II) or plumbous (Pb^{2+})	Thiocyanate (SCN^-)
Manganese(II) or manganous (Mn^{2+})	Group 17 anions (F^-, Cl^-, Br^-, I^-)
Mercury(I) or mercurous (Hg_2^{2+})	
Mercury(II) or mercuric (Hg^{2+})	***-2 Charge***
Nickel (Ni^{2+})	Carbonate (CO_3^{2-})
Tin(II) or stannous (Sn^{2+})	Chromate (CrO_4^{2-})
Zinc (Zn^{2+})	Dichromate ($Cr_2O_7^{2-}$)
Group 2 (Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+})	Hydrogen phosphate (HPO_4^{2-})
	Oxalate ($C_2O_4^{2-}$)
+3 Charge	Oxide (O^{2-})
Aluminum (Al^{3+})	Peroxide (O_2^{2-})
Chromium(III) or chromic (Cr^{3+})	Sulfate (SO_4^{2-})
Iron(III) or ferric (Fe^{3+})	Sulfite (SO_3^{2-})
	Sulfide (S^{2-})
+4 Charge	Thiosulfate ($S_2O_3^{2-}$)
Lead(IV) or plumbic (Pb^{4+})	
Tin(IV) or stannic (Sn^{4+})	***-3 Charge***
	Arsenate (AsO_4^{3-})
	Phosphate (PO_4^{3-})
	Phosphite (PO_3^{3-})
	Group 15 – nitride (N^{3-}), phosphide (P^{3-})

Summary of multivalent cations (metal cations with more than one possible charge):
Cu^+, Cu^{2+}; Hg^{2+}, Hg_2^{2+}; Co^{2+}, Co^{3+}; Cr^{2+}, Cr^{3+}; Fe^{2+}, Fe^{3+}; Mn^{2+}, Mn^{3+}; Pb^{2+}, Pb^{4+}; Sn^{2+}, Sn^{4+}
Manganese and some other metals can form several ions with different charges. You should know the ones listed.

Sulfite	Sulfate	Hydrogen sulfate
Phosphate	Dihydrogen Phosphate	Hydrogen Phosphate
Nitrite	Nitrate	Ammonium

Thiocyanate	Carbonate	Hydrogen carbonate
Borate	Chromate	Dichromate
Permanganate	Oxalate	Amide
Hydroxide	Cyanide	Acetate
Peroxide	Hypochlorite	Chlorite
Chlorate	Perchorate	Thiosulfate

HSO_4^-	SO_4^{2-}	SO_3^{2-}
HPO_4^{2-}	$H_2PO_4^-$	PO_4^{3-}
NH_4^+	NO_3^-	NO_2^-
HCO_3^-	CO_3^{2-}	NCS^- SCN^-
$Cr_2O_7^{2-}$	CrO_4^{2-}	BO_3^{3-}
NH_2^-	$C_2O_4^{2-}$	MnO_4^-

$C_2H_3O_2^-$ CH_3COO^-	CN^-	OH^-
ClO_2^-	ClO^-	O_2^{2-}
$S_2O_3^{2-}$	ClO_4^-	ClO_3^-

2. Solubility Rules

1. All common compounds of Group 1 and ammonium ions are soluble.
2. All nitrate, acetates, and chlorates are soluble.
3. All binary compounds of the halogens (other than F) are soluble, **except** those of Ag, Hg(I), and Pb.
4. All sulfates are soluble, **except** those of barium, strontium, calcium, and lead.
5. Sulfides and hydroxide are insoluble **except** for Ca, Ba, Sr, ammonium and the alkali metals.
6. Except for rule 1, carbonates, oxides, silicates, and phosphates are insoluble.

Note: You can apply the solubility rules to predict whether the product of a double replacement reaction will be a precipitate or not. If the compound is soluble, it will dissociate into free ions in solution. If the compound is insoluble, it will be a precipitate.

Study Suggestion: Prepare flash cards of the common ions with the formula on one side and the name on the other. Use these to quiz yourself on the names and formulas. To practice the solubility rules randomly pair a cation and an anion and predict if the resulting compound is soluble or not. Have the rules available as you are learning and then cover them up and quiz yourself.

3. Acids and Bases

Common Acids
Hydrochloric acid – HCl (aq)
Nitric acid – NHO_3 (aq)
Sulfuric acid – H_2SO_4 (aq)
Phosphoric acid – H_3PO_4 (aq)
Carbonic acid – H_2CO_3 (aq)
Acetic Acid (ethanoic acid) $HC_2H_3O_2$ (aq) or CH_3COOH (aq)

Strong Acids
HCl (aq)
HBr (aq)
HI (aq)
HNO_3 (aq)
$HClO_3$ (aq)

$HClO_4$ (aq)
H_2SO_4 (aq)

Common Bases
Sodium hydroxide – $NaOH$ (aq)
Potassium hydroxide – KOH (aq)
Calcium hydroxide – $Ca(OH)_2$ (aq)
Aqueous ammonia – NH_3 (aq)

Strong Bases
$LiOH$ (aq)
$NaOH$ (aq)
KOH (aq)
$RbOH$ (aq)
$Ca(OH)_2$ (aq)
$Sr(OH)_2$ (aq)
$Ba(OH)_2$ (aq)

Note: Strong acids and bases are those that dissociate completely in water.

Example: HCl (aq) → H^+ (aq) + Cl^-(aq)

$Ca(OH)_2$ (aq) → Ca^{2+} (aq) + $2OH^-$ (aq)

Weak acids and bases do not dissociate completely, and will be present as the compound in water.

Example: $HC_2H_3O_2$ (aq) ↔ H^+ (aq) + $C_2H_3O_2^-$ (aq)

NH_3 (aq) ↔ NH^+(aq) + OH^-(aq)

INFORMATION IN THE TABLE BELOW AND IN THE TABLES ON PAGES 3-5 MAY BE USEFUL IN ANSWERING THE QUESTIONS IN THIS SECTION OF THE EXAMINATION.

DO NOT DETACH FROM BOOK.

PERIODIC TABLE OF THE ELEMENTS

1 **H** 1.0079																	2 **He** 4.0026
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.811	6 **C** 12.011	7 **N** 14.007	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.179
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.453	18 **Ar** 39.948
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.938	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** 72.59	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** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 § (269)	111 § (272)	112 § (277)						

§Not yet named

*Lanthanide Series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** 237.05	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)

GO ON TO THE NEXT PAGE.

6. Significant Figures Rules

1.) all nonzero digits (1-9) ARE significant
2.) zeroes between non-zero digits are significant: the "oreo" cookie rule: 707 cannot be written without the zero there – otherwise the number becomes 77 – so that zero does indeed matter!
3.) IF the number has a decimal point in it – find the first non-zero digit in the number and ALL digits to the right of that ARE significant:

100.00 = **5 sig figs**: first non-zero digit in the number is a 1,

all digits to the right of that count as significant

0.002000 = **4 sig figs**: the first non-zero digit in the number is a 2,

All digits to the right of that count as significant

4.) Numbers without decimal point: icky numbers: If there is no decimal point in the number and the zero isn't sandwiched in the oreo cookie, then they are ambiguous, which means, we just don't know if they matter or not. These are icky numbers – because we don't always know how many sig figs there are. We can avoid this by using scientific notation!

3000 (this number can have 1, 2, 3 or even 4 sig figs!)

If we mean 1 sig fig then write the number as 3×10^3

If we mean 2 sig figs then write the number as 3.0×10^3

If we mean 3 sig figs then write the number as 3.00×10^3

If we mean 4 sig figs then write the number as 3.000×10^3

Notice all of these numbers still mean 3000, just with different numbers of sig figs!

<u>Multiplication and Division</u>: the final answer will have the same number of sig figs as the number in the problem with the fewest # of sig figs:

Example:

35 x 22.87663 = 800.68205 (what your calculator spits out)

Since 35 has only 2 sf while 22.87663 has 7 sf, the final answer can only have 2 sf

<u>Addition and Subtraction:</u> this rule has NOTHING to do with how many sig figs are in each of the individual numbers that are being added or subtracted. We don't give a hoot if one number has 3 sig figs and the other number has 20. Doesn't matter. What we do care about is the precision of the number – or how many digits there are past the decimal point.

The easiest way to follow sig figs in addition and subtraction is to line your numbers of vertically. If you don't have a number to add to, then you can't say what the number would be!

What the line is ultimately doing is showing us that when we add (or subtract) our final answer is as precise as the least precise number – or – our final answer will have as many decimal places as the number with the FEWEST decimals places in the addition or subtraction.

9 8 . 7 | 6 5 2 2 2 1 5 4

\- 9 2 . 1 |

6 . 6 | 6 5 2 2 2 1 5 4

The final answer would be 6.7 – notice the number after the line is a 6, sooo round it on up!

Notice that the least precise number only goes to the tenths place, (the number with the fewest decimal places), so the final answer can also only go to the tenths place!

<u>Mixed Operations</u>: Differing schools of thought – and will it ultimately matter? Some say, whenever you change major operations (e.g. switch between rules for operations, do addition then multiplication or subtraction and then multiplication), to round to the appropriate sig figs BEFORE continuing on. Others say, carry all the sig figs through in your calculator and round to the appropriate number at the end. That's all good, if you know how many you are supposed to have (which can be difficult to discern!)

For example:

$$\frac{(99.833 - 99.2)}{99.2} = ????$$

Well, let's do the subtraction first:

9 9 . 8 | 3 3

\- 9 9 . 2 |

0 . 6 | 3 3

EEK!!! Our answer can only go out to the tenths place and our answer from subtraction ONLY has 1 sf

Two choices: carry 0.633/99.2 through – realizing that your final answer can only have 1 sf

Or – round 0.633 to 1 sf, then divide by 99.2 – and your answer will have 1 sf

Let's see how doing both affects the final answer:

$$\frac{(0.633)}{99.2} = 0.00638 \qquad \frac{(0.6)}{99.2} = 0.0060$$

Well, no matter how you slice it, your answer can ONLY have 1 sf

0.00638 to 1 sf = **0.006** or 0.00605 to 1 sf = **0.006**

Which is why many people recommend that you round when you change operations. It's pretty clear from this operation: $\frac{(0.6)}{99.2} = ???$ That the final answer can only have 1 sig fig in it.

It's not so clear from this operation: $\frac{(0.633)}{99.2} = ???$ How many sig figs should be in your final answer. How you decide to do the mixed operations is up to you. But BE CAREFUL, because sig figs count ☺