Dexter Southfield Advanced Placement Chemistry
Summer Preparation
2021-2022

Greetings and Happy Summer!

According to the College Board, *"The AP Chemistry course is designed to be the equivalent of the general chemistry course generally taken during the first college year. There is an emphasis on chemical calculations and the mathematical formulation of principles, as well as advanced laboratory work. Students in an AP Chemistry course should expect to spend at least five hours a week in individual study outside of the classroom."* I am excited that you have accepted the challenge that an AP Chemistry course has to offer. I am also excited to work with you!

To ensure that all students in the AP Chemistry class are ready to partake in this high-paced, rigorous journey on the first day of school, the following summer assignment must be completed and turned in on the first day of class. The purposes of the assignment are to revisit chemical concepts learned in your first year chemistry class and expose you to the level of rigor demanded by the AP curriculum. This will allow us to focus our attention on the advanced chemistry topics and 16 suggested inquiry labs that will be tested on the AP exam in early May 2022.
Please work on these "5" Points and read the OPTIONAL paragraph.

1. A nice warm up:

Please make (and print) a document that contains the **name, picture, function,** and **unit it measures** of the following pieces of lab equipment:

Analytical balance	Pipette, graduated
Beaker	Pipette, beral-type
Buret	Pipette, volumetric
Calorimeter	Spectrophotometer
Cuvette	Volumetric flask
Graduated cylinder	Rubber Policeman

You should also include a **picture and a description** of the following additional pieces of equipment:

Crucible and cover	Stirring rod
Erlenmeyer flask	Watch glass
Funnel	Wire gauze
Mortar and pestle	Hot plate
Spatula	Test tubes
Scoopula	Clay triangle

2. A little bit more:

Read, take thoughtful notes, and do the exercises in the attached "The Ultimate Chemical Equations Handbook"- **Chapters 2, 3, and 4. (Separate paper)**

3. And essential for a quiz during week 1:

Please memorize the following polyatomic ions. Be able to name them and write them.
It is also essential that you memorize the solubility rules in the chart below.

Common Polyatomic Ions

+1		-1		-2		-3	
NH_4^+	ammonium	$C_2H_3O_2^-$	acetate	CO_3^{2-}	carbonate	PO_4^{3-}	phosphate
H_3O^+	hydronium	ClO^-	hypochlorite	CrO_4^{2-}	chromate	PO_3^{3-}	phosphite
		ClO_2^-	chlorite	$Cr_2O_7^{2-}$	dichromate		
		ClO_3^-	chlorate	SO_4^{2-}	sulfate		
		ClO_4^-	perchlorate	SO_3^{2-}	sulfite		
		CN^-	cyanide	O_2^{2-}	peroxide		
		NO_3^-	nitrate	$C_2O_4^{2-}$	oxalate		
		NO_2^-	nitrite				
		HCO_3^-	hydrogen carbonate (bicarbonate)				
		OH^-	hydroxide				
		MnO_4^-	permanganate				

TABLE 4.2 Solubility Rules for Common Ionic Compounds in Water at 25°C

Soluble Compounds	Insoluble Exceptions
Compounds containing alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+) and the ammonium ion (NH_4^+)	
Nitrates (NO_3^-), bicarbonates (HCO_3^-), and chlorates (ClO_3^-)	
Halides (Cl^-, Br^-, I^-)	Halides of Ag^+, Hg_2^{2+}, and Pb^{2+}
Sulfates (SO_4^{2-})	Sulfates of Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Compounds	**Soluble Exceptions**
Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), chromates (CrO_4^{2-}), sulfides (S^{2-})	Compounds containing alkali metal ions and the ammonium ion
Hydroxides (OH^-)	Compounds containing alkali metal ions and the Ba^{2+} ion

4. And key to review: Significant Figures

This topic might take a bit of study. Please see notes and problems to complete on page 17. There are rules and examples for addition, subtraction, multiplication, and division.

5. You will survive: Videos

Please watch these videos and take notes. Writing equations is foundational for AP Chemistry.

https://www.youtube.com/watch?v=kme6ditQieY
(Types of Chemical reactions by Jon Bergmann)

https://www.youtube.com/watch?v=uFl2tHOSqBo
(Net ionic equations by Jon Bergmann)

Optional:

I highly recommend taking out some of your Chemistry notes from sophomore year. Were there topics (such as gas laws or stoichiometry) that were challenging? Now is a good time to review!

My favorite book: https://www.amazon.com/Homework-Helpers-Chemistry-Greg-Curran/dp/1601631634 is very affordable and students have read it over the summer as a guide (like having a tutor!)

I also like:
http://www.phschool.com/webcodes10/index.cfm?fuseaction=home.gotoWebCode&wcprefix=cdk&wcsuffix=0000
Since the website grades 10 question quizzes on all of the topics covered in a general chemistry course! It is free and fun to practice.

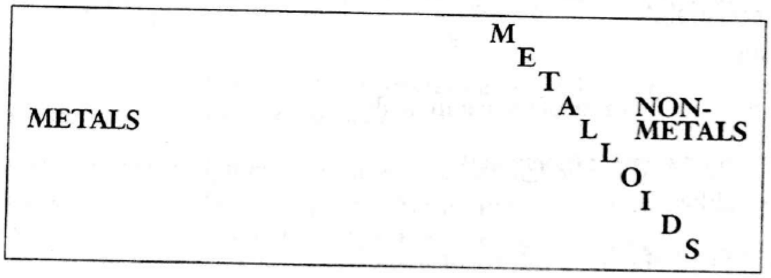
Chapter 2
Simple Inorganic Formulas and Nomenclature

Compounds consisting of two different elements in various ratios are considered to be *binary compounds.* Binary compounds usually end in the suffix "ide." There are two types of binary compounds—binary molecules and binary salts. A binary molecule consists of two nonmetals bonded via covalent bonding. A binary salt consists of a metal and a nonmetal exhibiting ionic bonding.

General Rules

A. Binary Molecules (Nonmetal + Nonmetal) i.e., CO_2 or N_2O_3

Molecules are formed when two nonmetals or metalloids combine and prefixes must be used to designate the number of atoms of each element present in one molecule. Nonmetals are found just to the right of the zigzag line on the periodic table. Metalloids are near the zigzag line and may have some properties of metals and other properties of nonmetals.



Prefixes are used to designate the number of atoms of each element present in the formula of a binary compound. The prefix *mono* is **never** used in front of the first element (standard convention). If there is only one atom, the mono is assumed.

1 = mono	4 = tetra	7 = hepta	10 = deca
2 = di	5 = penta	8 = octa	11 = undeca
3 = tri	6 = hexa	9 = nona	12 = dodeca

Name the following binary molecules — CO_2 and N_2O_3

To determine the first word in the name of the compound:

1. Give the prefix designating the number of atoms of the first element present. Remember, *mono* is never used (by standard convention) for the first element.

   CO_2: No prefix for C — N_2O_3: **di**

2. Name the first element.

   CO_2: **carbon** — N_2O_3: **dinitrogen**

To determine the second word in the compound's name:

3. Give the prefix designating the number of atoms of the second element present.
   CO_2: carbon **di** N_2O_3: dinitrogen **tri**
4. Name the root of the second element. *Note:* The root is the base name that designates the element.
   CO_2: carbon **diox** N_2O_3: dinitrogen **triox**
5. Add the suffix *–ide* to the root of the second element.
   CO_2: carbon diox**ide** (official name) N_2O_3: dinitrogen triox**ide** (official name)

B. Binary Salts (Metal + Nonmetal) i.e., $CaCl_2$

Prefixes giving the number of atoms of each element present are *never* used to name an ionic salt. Salts exhibit ionic bonding between a metal and a nonmetal, while molecular substances exhibit covalent bonding between two nonmetals.

Name the following binary salt — $CaCl_2$

By convention, the metal is written before the nonmetal. To identify the first word in the name:

1. Name the first element (metal).
   $CaCl_2$: **calcium**

To determine the second word in the name of the compound:

2. Name the root of the second element (nonmetal).
   $CaCl_2$: calcium **chlor**
3. Add the suffix *–ide* to the root of the second element.
   $CaCl_2$: calcium chlor**ide**

Exercise 2–1: In column 1, classify each of the following compounds as binary molecules (M) or binary ionic salts (I). Then in column 2, use the rules to name each binary compound.

1. CaF_2	____	____________	10. SrI_2	____	____________
2. P_4O_{10}	____	____________	11. CO	____	____________
3. K_2S	____	____________	12. Cs_2Po	____	____________
4. NaH	____	____________	13. $ZnAt_2$	____	____________
5. Al_2Se_3	____	____________	14. P_4S_3	____	____________
6. N_2O	____	____________	15. AgCl	____	____________
7. O_2F	____	____________	16. Na_3N	____	____________
8. SBr_6	____	____________	17. Mg_3P_2	____	____________
9. Li_2Te	____	____________	18. XeF_6	____	____________

Chapter 3
Oxidation Numbers: Anions and Cations

Metals with Variable Charges (Oxidation Numbers)

A number of metallic elements can form compounds in which the metal ions (cations) may have different charges. These charges are known as oxidation numbers and are sometimes referred to as valences. The transition metals in the middle of the periodic table have variable oxidation numbers as do many of the representative elements in groups 13–16 in the periodic table. Cations with variable oxidation numbers use a Roman numeral enclosed in parentheses to designate the charge on the metal ion. This naming system is called the Stock System. For example, the oxidation number of iron in the following two compounds cannot be the same: $FeCl_2$ and $FeCl_3$. Calling both of these compounds iron chloride would only lead to confusion. The Stock System is used to differentiate between ions that have two or more possible charges. $FeCl_2$ is known as iron(II) chloride and $FeCl_3$ is officially called iron(III) chloride. The Roman numeral represents the charge on the metal cation and does *not* represent the number of atoms of the element present. To name these types of ionic compounds, the oxidation numbers of all the elements present must be known.

Here are some simple rules that should help in the determination of the oxidation numbers of metallic ions (cations) from the formulas of their compounds.

1. The oxidation number of any **element** in its free state (uncombined with other elements) is zero, e.g., Fe in a bar of iron is zero. O_2 and N_2 in the Earth's atmosphere both have oxidation numbers of zero. When an element has equal numbers of protons and electrons, its overall charge is zero.
2. The oxidation number of **alkali metals** in a compound is always 1+, e.g., Li^+, Na^+, K^+, etc.
3. The oxidation number of **alkaline earth metals** in a compound is always 2+, e.g., Mg^{2+}, Ca^{2+}, Sr^{2+}, etc.
4. **Fluorine** is always assigned an oxidation number of 1^- in a compound, e.g., F^-.
5. The oxidation number of **oxygen** is almost always 2^- in a compound. Exceptions to this rule would be peroxides, O_2^{2-} where the oxidation number of each oxygen is 1^-, and superoxides, O_2^- where the oxidation number of each oxygen is $\frac{1}{2}^-$. Neither peroxides nor superoxides are common. Peroxides are only known to form compounds with the elements in the first two columns of the periodic table, e.g., H_2O_2, Na_2O_2, CaO_2, etc. Potassium, rubidium, and cesium are the only elements that form superoxides, e.g., KO_2. *Note:* The name superoxide may also be called superperoxide.
6. In covalent compounds with nonmetals, **hydrogen** is assigned an oxidation number of 1+, e.g., HCl, H_2O, NH_3, CH_4. The exception to this rule is when hydrogen combines with a metal to form a **hydride**. Under these conditions, which are rare, hydrogen is assigned an oxidation number of 1^-, e.g., NaH.
7. In **metallic halides** the halogen (F, Cl, Br, I, At) always has an oxidation number equal to 1^-.
8. Sulfide, selenide, telluride, and polonide are always 2^- in binary salts.
9. Nitrides, phosphides, and arsenides are always 3^- in binary salts.
10. All other oxidation numbers are assigned so that the sum of the oxidation numbers of each element equals the net charge on the molecule or polyatomic ion. In a neutral compound, the sum of the positive and negative charges must always equal zero.

Example

Determine the oxidation number of the underlined element: $K\underline{Mn}O_4$. Since K is an alkali metal, its charge must be 1+. Oxygen is 2– but there are four of them, therefore, 4 times 2– equals 8–. If 1+ and 8– are added together, we get 7–. In order for the compound to be neutral, the Mn must be 7+.

Algebraically, $(1+) + (x) + 4(2-) = 0 \quad \therefore \quad x = 7+$

Other Examples

$\underline{N}H_4^+$: The sum of the charges on this polyatomic ion must equal 1+. Since hydrogen has a 1+ charge and there are four hydrogen atoms, the nitrogen must be 3– because (3–) + (4+) = 1+!

$K_2\underline{Cr}_2O_7$: Potassium is 2 times 1+ = 2+, and oxygen is 7 times 2– = 14–. (14–) + (2+) = 12–. Since there are two chromium atoms and the compound is neutral overall, the charge on the two chromium atoms must be equal to 12+ and each chromium atom must have a charge of 6+ (since 12+/2 = 6+).

Algebraically, $2(1+) + (2x) + 7(2-) = 0 \quad \therefore \quad x = 6+$

$\underline{O}_2$: This is an element in its free state, so the oxidation number must be zero.

Note: Ions written alone, such as peroxide, must be written with a charge on them, e.g., O_2^{2-}. In a compound, the charges on individual atoms or ions are not shown.

Exercise 3–1: Determine the oxidation number of each underlined element.

1. $K_2\underline{S}$
2. $Na\underline{Cl}O_4$
3. $\underline{Br}Cl$
4. $Li_2\underline{C}O_3$
5. $\underline{O}F_2$
6. $\underline{S}_8$
7. $\underline{Mg}$
8. $K_2\underline{W}_4O_{13}$
9. $Mg(\underline{B}F_4)_2$
10. $\underline{Au}_2O_3$
11. $\underline{C}_{60}$
12. $\underline{Zr}O_2$
13. $\underline{Nb}OF_6^{3-}$
14. $Al_2(\underline{Cr}O_4)_3$
15. $Cs_2\underline{Te}F_8$

Remember, free elements, no matter how complex the molecule, have an oxidation number (valence or charge) equal to zero. The following are diatomic or polyatomic elements in nature which must be committed to memory. These elements exist as neutral molecules in nature!

Polyatomic Elements

Hydrogen, H_2	Bromine, Br_2
Nitrogen, N_2	Iodine, I_2
Oxygen, O_2	Ozone, O_3
Fluorine, F_2	Phosphorus, P_4
Chlorine, Cl_2	Sulfur, S_8

Most common forms of buckminsterfullerenes (buckyballs): C_{60} and C_{70}

Representative Elements (s- or p-block) Cations and Anions

Charges can be determined by position (family) on the Periodic Table. Cations (+ ions) come from metals that lose electrons (oxidation) in order to become isoelectronic with a noble gas. Anions (– ions) come from nonmetals that gain electrons (reduction) to become isoelectronic with a noble gas.

Oxidation Numbers (Valence) of Representative Element Cations and Anions							
1+ Alkali metals	2+ Alkaline earth metals	3+		4–	3– Nitrogen family	2– Oxygen family	1– Halogens
Lithium Sodium Potassium Rubidium Cesium Francium Hydrogen	Magnesium Calcium Strontium Barium Radium Beryllium	Aluminum Boron		Carbide	Nitride Phosphide Arsenide	Oxide Sulfide Selenide Telluride Polonide	Fluoride Chloride Bromide Iodide Astatide

More on Metallic Elements with Variable Oxidation Numbers

Transition metals, representative metals with *p* and *d* sublevels, and the inner transition metals typically have more than one oxidation state in compounds. Electrons for these metallic elements are lost (oxidized) from their outermost energy levels in the following order: *p*, *s*, *d*. Such elements are ***not*** isoelectronic with a noble gas when the outermost (valence) electrons are lost and if enough energy is available, will begin to lose *d* level electrons.

Example 1: A neutral vanadium atom has an electron configuration of [Ar] $4s^2$ $3d^3$. The outermost electrons are always lost first, therefore, vanadium will lose its $4s^2$ electrons and form the vanadium(II) ion, V^{2+}. With additional energy, the V^{2+} cation can lose its $3d^3$ electrons in order, forming vanadium(III), V^{3+}, vanadium(IV), V^{4+}, and vanadium(V), V^{5+} cations.

Example 2: The electron configuration for an atom of Fe is [Ar] $4s^2$ $3d^6$. The first cation that forms when the $4s^2$ electrons are lost is the iron(II) ion, Fe^{2+}. Additional energy will cause the iron(II) ion to lose one of its 3d electrons to form the iron(III) ion, Fe^{3+}. The remaining *d* electrons are all spinning in the same direction and the energy required to oxidize them is greater than normally encountered in an ordinary chemical reaction. The repulsive forces between the only two paired electrons in the 3d sublevel make the formation of the iron(III) ion relatively easy.

Example 3: The electronic configuration of a neutral lead atom is [Xe] $6s^2$ $4f^{14}$ $5d^{10}$ $6p^2$. The two common oxidation numbers of lead are lead(II) when the two $6p^2$ electrons are lost and lead(IV) when the two $6s^2$ electrons are also oxidized. Tin behaves in a similar manner when it forms tin(II) and tin(IV) cations. Bismuth with an electron configuration of [Xe] $6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$, forms bismuth(III) and bismuth(V) ions.

Inner transition elements are sometimes called by such names as the lanthanides, actinides, rare earth elements, and the transuranium elements. All of these elements are quite rare, and many of the elements beyond uranium (the transuranium elements) exist for only short periods of time. Reactions involving such elements are seldom encountered in a beginning chemistry course and there is little need to pursue this topic in any detail. Two inner transition elements worth mentioning are uranium (U^{3+}, U^{4+}, and U^{5+}) and cerium (Ce^{3+} and Ce^{4+}).

Both inner transition and transition elements are known for their variable oxidation numbers. The most common oxidation number for transition elements is 2+. The *d* sublevel in transition elements is responsible for the various oxidation numbers that result. Incomplete *d* sublevels are also responsible for the many colorful transition compounds that are known to exist. Complete *d* sublevels in cations of silver and zinc result in white compounds.

Summary of Cations with Variable Oxidation Numbers—Stock System

1+, 2+	copper(I), Cu^{+}; copper(II), Cu^{2+};
	mercury(I)*, Hg_2^{2+}; mercury(II), Hg^{2+}
	**Note: mercury(I) actually exists as a diatomic ion and is written as* Hg_2^{2+} *and not* Hg^{+}.
1+, 3+	gold(I), Au^{+}; gold(III), Au^{3+};
	indium(I), In^{+}; indium(III), In^{3+};
	thallium(I), Tl^{+}; thallium(III), Tl^{3+}
2+, 3+	chromium(II), Cr^{2+}; chromium(III), Cr^{3+};
	cobalt(II), Co^{2+}; cobalt(III), Co^{3+};
	iron(II), Fe^{2+}; iron(III), Fe^{3+};
	manganese(II), Mn^{2+}; manganese(III), Mn^{3+}
2+, 4+	lead(II), Pb^{2+}; lead(IV), Pb^{4+};
	platinum(II), Pt^{2+}; platinum(IV), Pt^{4+};
	tin(II), Sn^{2+}; tin(IV), Sn^{4+};
	zirconium(II), Zr^{2+}; zirconium(IV), Zr^{4+}
3+, 4+	cerium(III), Ce^{3+}; cerium(IV), Ce^{4+}
3+, 5+	antimony(III), Sb^{3+}; antimony(V), Sb^{5+};
	arsenic(III), As^{3+}; arsenic(V), As^{5+};
	bismuth(III), Bi^{3+}; bismuth(V), Bi^{5+};
	phosphorus(III), P^{3+}; phosphorus(V), P^{5+}
2+, 3+, 4+	iridium(II), Ir^{2+}; iridium(III), Ir^{3+}; iridium(IV), Ir^{4+};
	titanium(II), Ti^{2+}; titanium(III), Ti^{3+}; titanium(IV), Ti^{4+}
2+, 4+, 5+	tungsten(II), W^{2+}; tungsten(IV), W^{4+}; tungsten(V), W^{5+}
3+, 4+, 5+	uranium(III), U^{3+}; uranium(IV), U^{4+}; uranium(V), U^{5+}
2+, 3+, 4+, 5+	vanadium(II), V^{2+}; vanadium(III), V^{3+};
	vanadium(IV), V^{4+}; vanadium(V), V^{5+}

Note: When reading the name of an ion such as Pb^{2+}, *the ion is read in English as the "lead two ion."*

Special Metallic Cations

The following transition metal cations do not exhibit variable oxidation numbers and are normally written without Roman numerals:

cadmium, Cd^{2+} silver, Ag^{+} zinc, Zn^{2+}

Nickel, on the other hand, has variable oxidation numbers, and even though it almost always appears as the nickel(II) ion, Ni^{2+}, the Roman numeral must be written.

The ions of the representative elements gallium, germanium, and indium do not have variable oxidation numbers, but are written with Roman numerals:

gallium(III), Ga^{3+} germanium(IV), Ge^{4+} indium(III), In^{3+}

Polyatomic Ions

The term polyatomic ion is used to describe a group of atoms that behave as a single ion. The bonding within a polyatomic ion is covalent, but because there is always an excess or shortage of electrons when compared to the number of protons present, an ion results. A common polyatomic positive ion (cation) is the ammonium ion, NH_4^+. A common polyatomic negative ion (anion) is the sulfate ion, SO_4^{2-}.

Remember that polyatomic ions stay together as a group. The ammonium ion is always written as NH_4^+ and *never* as $N^{3-} + 4H^+$ or H_4^+ or H_4^{4+}. If two or more of the same polyatomic ions are needed within a compound in order to reach electrical neutrality, the polyatomic group is enclosed in parentheses. For example, ammonium sulfate is written as $(NH_4)_2 SO_4$. The compound consists of two ammonium ions and one sulfate ion. The letters are read as "N, H, four taken twice, S, O, four."

Polyatomic ions must be memorized! There is no simple way to learn all of these ions but it is helpful to realize that some of them come in related pairs. For example, sulfate, SO_4^{2-}, and sulfite, SO_3^{2-} share the same charge and include the same elements, S and O, but they differ in their number of oxygen atoms. Notice that the *–ate* form has one more oxygen atom that the *–ite* form; in other words sulf<u>ate</u> "ate" one more O than sulf<u>ite</u>. There are several of these pairs, so if you know nitrate is NO_3^-, then it's easy to deduce that nitrite is NO_2^-. Chlorate is ClO_3^- and chlorite is ClO_2^-.

Another helpful tip is to observe patterns in the *–ate* formulas and their relationship to the periodic table. Notice that all of the *–ate* ions on the outside of the bold line have three oxygen atoms and the *–ate* ions on the inside of the bold line have four oxygen atoms.

	CO_3^{2-}	NO_3^-		
MnO_4^-		PO_4^{3-}	SO_4^{2-}	ClO_3^-
CrO_4^{2-}			SeO_4^{2-}	BrO_3^-
				IO_3^-

Common Polyatomic Ions

Anions

1–

acetate, CH_3COO^-
amide, NH_2^-
azide, N_3^-
benzoate, $C_6H_5COO^-$
bromate, BrO_3^-
chlorate, ClO_3^-
chlorite, ClO_2^-
cyanate, OCN^-
cyanide, CN^-
dihydrogen phosphate, $H_2PO_4^-$
formate, $HCOO^-$
hydrogen carbonate, HCO_3^-
(bicarbonate)
hydrogen sulfate, HSO_4^-
(bisulfate)
hydrogen sulfide, HS^-
(bisulfide or hydrosulfide)
hydroxide, OH^-
(called hydroxyl when aqueous)
hypochlorite, ClO^-
iodate, IO_3^-
nitrate, NO_3^-
nitrite, NO_2^-
perchlorate, ClO_4^-
permanganate, MnO_4^-
thiocyanate, SCN^-
(thiocyanato)
triiodide, I_3^-
vanadate, VO_3^-

2–

carbide, C_2^{2-}
(saltlike)
carbonate, CO_3^{2-}
chromate, CrO_4^{2-}
dichromate, $Cr_2O_7^{2-}$
imide, NH^{2-}
manganate, MnO_4^{2-}
metasilicate, SiO_3^{2-}
monohydrogen phosphate, HPO_4^{2-}
oxalate, $C_2O_4^{2-}$
peroxide, O_2^{2-}
peroxydisulfate, $S_2O_8^{2-}$
phthalate, $C_8H_4O_4^{2-}$
polysulfide, S_x^{2-}
selenate, SeO_4^{2-}
sulfate, SO_4^{2-}
sulfite, SO_3^{2-}
tartrate, $C_4H_4O_6^{2-}$
tellurate, TeO_4^{2-}
tetraborate, $B_4O_7^{2-}$
thiosulfate, $S_2O_3^{2-}$
tungstate, WO_4^{2-}
zincate, ZnO_2^{2-}

3–

aluminate, AlO_3^{3-}
arsenate, AsO_4^{3-}
borate, BO_3^{3-}
citrate, $C_6H_5O_7^{3-}$
phosphate, PO_4^{3-}

4–

orthosilicate, SiO_4^{4-}
pyrophosphate, $P_2O_7^{4-}$

5–

tripolyphosphate, $P_3O_{10}^{5-}$

Cations

1+

ammonium, NH_4^+
hydronium, H_3O^+

Exercise 3–2: Name the following substances.

1. $FeSO_3$ ____________________
2. $Cu(NO_3)_2$ ____________________
3. Hg_2Cl_2 ____________________
4. AgBr ____________________
5. $KClO_3$ ____________________
6. $MgCO_3$ ____________________
7. BaO_2 ____________________
8. KO_2 ____________________
9. SnO_2 ____________________
10. $Pb(OH)_2$ ____________________
11. $Ni_3(PO_4)_2$ ____________________
12. $CuCH_3COO$ ____________________
13. N_2O_4 ____________________
14. Rb_3P ____________________
15. S_8 ____________________
16. Fe_2O_3 ____________________
17. $(NH_4)_2SO_3$ ____________________
18. $Ca(MnO_4)_2$ ____________________
19. PF_5 ____________________
20. LiH ____________________

Exercise 3–3: Write formulas for the following substances.

1. vanadium(V) oxide ____________________
2. dihydrogen monoxide ____________________
3. ammonium oxalate ____________________
4. polonium(VI) thiocyanate ____________________
5. tetraphosphorus decaoxide ____________________
6. zinc hydroxide ____________________
7. potassium cyanide ____________________
8. cesium tartrate ____________________
9. oxygen molecule ____________________
10. mercury(II) acetate ____________________
11. silver chromate ____________________
12. tin(II) carbonate ____________________
13. sodium hydrogen carbonate ____________________
14. manganese(VII) oxide ____________________
15. copper(II) dihydrogen phosphate ____________________
16. francium dichromate ____________________
17. calcium carbide ____________________
18. mercury(I) nitrate ____________________
19. cerium(IV) benzoate ____________________
20. potassium hydrogen phthalate ____________________

Chapter 4
Ternary Nomenclature: Acids and Salts

The halogens, with their variable oxidation numbers, allow for a great variety of compounds. The problem arises on how these compounds should be named. For example, chlorine is found with a different oxidation state in each of the following five compounds:

$HClO_4$ (Cl = 7+)

$HClO_3$ (Cl = 5+)

$HClO_2$ (Cl = 3+)

HClO (Cl = 1+)

HCl (Cl = 1–)

A good way to learn ternary nomenclature is to start with a certain "home base" polyatomic ion. This is the polyatomic ion ending with the suffix *–ate* (see page 16). Remembering that salts are named by adding the name of the metallic ion (cation) to the nonmetallic polyatomic ion (anion), the following rules apply:

Number of Oxygen Atoms (Compared to Home Base)	Polyatomic Ion Name		Acid Name (H^+ Combined with Polyatomic Ion)	
Plus One Oxygen Atom	ClO_4^-	*per*chlor*ate*	$HClO_4$	*per*chlor*ic* acid
Home Base	ClO_3^-	chlor*ate*	$HClO_3$	chlor*ic* acid
Minus One Oxygen Atom	ClO_2^-	chlor*ite*	$HClO_2$	chlor*ous* acid
Minus Two Oxygen Atoms	ClO^-	*hypo*chlor*ite*	HClO	*hypo*chlor*ous* acid
No Oxygen Atoms	Cl^-	chlor*ide*	HCl*	*hydro*chlor*ic* acid

*Binary compounds containing hydrogen and a nonmetallic ion, such as hydrogen chloride, form acids when dissolved in water. The name of the resulting acid is derived by adding the prefix *hydro-* to the root name followed by the suffix *-ic* and the word acid. Thus, HCl gas is called hydrogen chloride (hydrogen monochloride), but is known as hydrochloric acid in aqueous solution.

Common Binary Acids

Formula	Name	Anion
HF(aq)	*hydro*fluor*ic* acid	F^-, fluor*ide* ion
HCl(aq)	*hydro*chlor*ic* acid	Cl^-, chlor*ide* ion
HBr(aq)	*hydro*brom*ic* acid	Br^-, brom*ide* ion
HI(aq)	*hydro*iod*ic* acid	I^-, iod*ide* ion
H_2S(aq)	*hydro*sulfur*ic* acid	S^{2-}, sulf*ide ion*

Many common acids contain only oxygen, hydrogen, and a nonmetallic ion or a polyatomic ion. Such acids are called ***oxyacids***. The suffixes *-ous* and *-ic* give the oxidation state of the atom bonded to the oxygen and the hydrogen. The *-ous* suffix always indicates the lower oxidation state and *-ic* the higher.

Common Oxyacids

Formula	Name	Anion	
$HClO_4$	*per*chlor*ic* acid	ClO_4^-	*per*chlor*ate*
$HClO_3$	chlor*ic* acid	ClO_3^-	chlor*ate*
$HClO_2$	chlor*ous* acid	ClO_2^-	chlor*ite*
$HClO$	*hypo*chlor*ous* acid	ClO^-	*hypo*chlor*ite*
HNO_3	nitr*ic* acid	NO_3^-	nitr*ate*
HNO_2	nitr*ous* acid	NO_2^-	nitr*ite*
H_2SO_4	sulfur*ic* acid	SO_4^{2-}	sulf*ate*
H_2SO_3	sulfur*ous* acid	SO_3^{2-}	sulf*ite*
CH_3COOH or $HC_2H_3O_2$	acet*ic* acid	CH_3COO^- or $C_2H_3O_2^-$	acet*ate*
H_2CO_3	carbon*ic* acid	CO_3^{2-}	carbon*ate*
$H_2C_2O_4$	oxal*ic* acid	$C_2O_4^{2-}$	oxal*ate*
H_3PO_4	phosphor*ic* acid	PO_4^{3-}	phosph*ate*

Exercise 4–1: Name the following compounds.

1. HIO_3
2. $NaBrO_2$
3. $Ca_3(PO_4)_2$
4. HIO_4
5. $Fe(IO_2)_3$
6. HAt(aq)
7. C_6H_5COOH
8. $Hg_2(IO)_2$
9. H_3PO_3
10. NH_4BrO_3

Exercise 4–2: Write formulas for the following compounds.

1. tartaric acid
2. calcium hypochlorite
3. hydrotelluric acid
4. copper(II) nitrite
5. carbonic acid
6. hypoiodous acid
7. cyanic acid
8. phthalic acid
9. tin(IV) chromate
10. selenic acid

DO YOU KNOW YOUR ACIDS?

-IC from *-ATE* *-OUS* from *-ITE* *HYDRO-, -IC, -IDE*

Exercise 4–3: Complete the following table.

Name of Acid	Formula of Acid	Name of Anion
*hydro*chlor*ic* acid	HCl	chlor*ide*
sulfur*ic* acid	H_2SO_4	sulf*ate*
	HI	
		sulf*ite*
chlor*ous* acid		
		nitr*ate*
	CH_3COOH or $HC_2H_3O_2$	
*hydro*brom*ic* acid		
		sulf*ide*
	HNO_2	
chrom*ic* acid		
		phosph*ate*

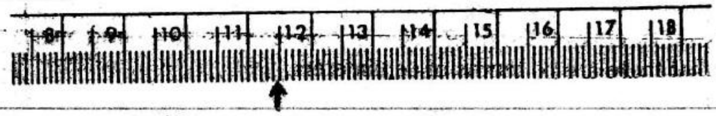
Significant Figures

The purpose of this worksheet is to familiarize you with the rules regarding significant figures, and to provide practice in interpreting data and manipulating data using the proper number of significant figures.

INTRODUCTION

WHAT ARE SIGNIFICANT FIGURES (SIGNIFICANT DIGITS)?

The digits known with certainty in any measured quantity (e.g. mass, length, volume) and the first estimated digit are known as significant figures, or significant digits.



WHY IS BEING CAREFUL TO MAINTAIN THE PROPER NUMBER OF SIGNIFICANT DIGITS WHEN RECORDING VALUES FOR MEASUREMENTS IMPORTANT?

The key to the answer to this question is in the word "significant". When taking data, any numbers tacked on to the end of a measurement in which the digits known with certainty and one estimated digit have already been recorded are meaningless, or insignificant. For example, in the ruler shown above, recording the value as 11.472 cm would be a misrepresentation of the data. There is no way, based on the scale of the ruler, to determine the value for the digit in the thousandths place. Using the proper number of significant digits in measurements allows recording of data as precisely as possible without misrepresenting the data. Furthermore, any conclusions based upon digits recorded unreliably have a high probability of being wrong. The conclusions are certainly without scientific merit.

Based on what you have read, answer the following two questions:

WHY IS MAINTAINING THE PROPER NUMBER OF SIGNIFICANT FIGURES WHEN MANIPULATING DATA THROUGH CALCULATIONS (e.g. MULTIPLICATION, DIVISION) IMPORTANT?

Answer:

THERE ARE SPECIFIC RULES FOR WRITING DOWN VALUES OF MEASUREMENTS SUCH THAT ANYONE FAMILIAR WITH THE RULES WILL KNOW HOW PRECISELY THE MEASUREMENT WAS MADE (e.g. which digit is the estimated one). WHY WAS ESTABLISHING A SET OF RULES IMPORTANT, AND WHY IS LEARNING THIS SET OF RULES IMPORTANT?

Answers:

SIGNIFICANT FIGURE RULES

Do you know what significant figures are and why they are important? If not, go back to the introduction. If yes, you are ready to continue......

I. RULE FOR TAKING MEASUREMENTS IN THE PROPER NUMBER OF SIGNIFICANT DIGITS

The rule for taking values for measurements in the proper number of significant digits comes directly from the definition of significant digits. Record all of the digits you know for certain, plus one estimated digit.

II. RULES FOR WRITING VALUES IN THE PROPER NUMBER OF SIGNIFICANT DIGITS

Values for measurements and derived values (values calculated using measurements, such as velocity, which is derived from distance and time) must be written according to specific guidelines such that anyone reading the values will know how precisely they were measured. The best way to learn how to write a value for a measurement or calculation in the proper number of significant figures is to first learn how to determine which figures are significant in measurements recorded by someone else. Several rules apply.

A. Nonzero digits

All nonzero numbers in a value properly recorded are significant. For example, 2.3 m has two significant figures.

PRACTICE How many significant figures does each of the following values have?

45 m	_______	0.956 s	_______
1.345 m	_______	2.5164 mm	_______
342,987 m	_______	8 m	_______

B. Zeros--three possibilities (while going through the following, try to think of <u>why</u> each rule makes sense). This is the only way you will remember it long term!

1. 1st possibility--zeros to the left of the last nonzero digit

   Zeros to the left of the last nonzero digit in a measurement or value properly recorded are <u>never</u> significant. For example, 0.0051 m and 0.0000067 m both have only two significant figures.

PRACTICE How many significant figures does each of the following measurements have?

a.	0.068765 m	_______	d.	0.0956 s	_______
b.	00487 m	_______	e.	8 mm	_______
c.	0.00000049 m	_______	f.	0.008 m	_______

<u>Review Question</u>: What does saying that a digit is significant in a recorded measurement mean?

Now, for the BIG question...WHY does it make sense that zeros to the left of the last nonzero digit in a properly recorded value would be considered insignificant? Hint: the two lengths in e and f above are identical. However, one is measured in meters, the other in millimeters.

YOUR ANSWER TO BIG Q #1:

2. 2nd possibility--zeros are in between two nonzero digits

   Zeros in between two nonzero digits in a value or measurement properly recorded are <u>always</u> significant. For example, 10001 and 3.4075 both have 5 significant digits.

PRACTICE How many significant figures does each of the following measurements have?

a.	908 m	_______	d.	0.705061 s	_______
b.	2.0057 m	_______	e.	4.078 mm	_______
c.	10000001 m	_______	f.	0.80808 m	_______

Now, for the second BIG question...WHY does it make sense that zeros in between nonzero digits in a measurement properly recorded would always be considered significant? Hint: consider the alternative, that they are not significant, and try to make sense out of it.

YOUR ANSWER TO BIG Q #2:

3. 3rd possibility--zeros are to the right of the last nonzero digit

Zeros to the right of the last nonzero digit in the value of a measurement or calculation properly recorded may or may not be significant.

-Zeros to the right of a nonzero digits are always significant if a decimal point appears in the number. For example, 300. g and 4.00 g both have three significant figures.

-Zeros to the right of a nonzero digit in a measurement or calculation properly recorded are never significant if no decimal point appears in the number. For example, 300 and 3,000,000 both have only one significant figure.

PRACTICE How many significant figures does each of the following measurements have?

a.	90000 m	_______	g.	7800 s	_______
b.	9.000 m	_______	h.	230. mm	_______
c.	357,000 m	_______	i.	56,450 m	_______
d.	56.980 m	_______	j.	9.80000 sec	_______
e.	100. m	_______	k.	4.078 mm	_______
f.	10000000 m	_______	l.	7800.000 m	_______

Now, for the third BIG question...WHY does it make sense that zeros to the right of the last nonzero digit would be considered insignificant in a measurement or calculation that does not have a decimal point?

YOUR ANSWER TO BIG Q #3:

C. Scientific notation and significant figures

Scientific notation is commonly used in chemistry because huge (astronomical might be a better word) and small infinitesimal might be a better word) numbers are commonly used. For example, There are 6.02×10^{23} atoms of carbon in 12.0 grams of carbon, and one atom of carbon has an average mass of 1.99×10^{-23} g. **For a review of how to write numbers in scientific notation, see sheet that accompanies this worksheet. In terms of significant figures, the digits not associated with the exponent are all significant; numbers associated with the exponent are not significant. For example 3.2×10^{23} has two significant figures, 4.00×10^{23} has three significant figures, and 5.679×10^{23} has four significant figures.

PRACTICE How many significant figures does each of the following have?

a.	5×10^{2} m	_______	c.	9.2005×10^{-23} g	_______
b.	3.02×10^{24} atoms	_______	d.	6.000×10^{23} atoms	_______

Now, for the fourth BIG question...WHY does the basic rule for writing numbers in scientific notation (see sheet that accompanies this one) make the determination of the number of significant digits in a value properly recorded in scientific notation easy to determine? Hint: think about the role of the decimal point.

YOUR ANSWER TO BIG Q #4:

D. PRACTICE--USING ALL OF THE ABOVE RULES

Here comes the tricky part. Determine the number of significant figures in each of the following properly written values. Using all of the above rules (in some cases, more than one rule may apply).

a.	0.00506 m	_______	g.	7500.0 s	_______
b.	500300 m	_______	h.	230 mm	_______

c.	4.000×10^4 m	______	i.	506,450 m	______
d.	56.0980 m	______	j.	9×10^{-28} s	______
e.	7000 m	______	k.	0.01010 mm	______
f.	0.00800 m	______	l.	9.02×10^{23} atoms	______

III. MAINTAINING THE PROPER NUMBER OF SIGNIFICANT FIGURES DURING CALCULATIONS

Do you remember why it is important to maintain the proper number of significant figures on calculated values? If not, go back to the beginning.

A. RULES FOR ROUNDING OFF CALCULATED VALUES

Before doing actual calculations it is important to know how to round numbers off to the proper number of significant figures. Three rules apply.

1. Rounding rule #1: Round to the value with the proper number of significant figures that is closest to the calculated value . (Determining how many digits the calculated value should have, or to which place the value should be rounded, will be explained below. For example, 0.367 m/s, rounded off to two significant figures, would be 0.37 m/s rather than 0.36 m/s, since 0.37 m/s is closer to 0.367 m/s than is .36 m/s.

Other examples, rounding off to 3 significant figures:

32.53 m/s ---> 32.5 m/s
32.58 m/s ---> 32.6 m/s
4563 m/s ---> 4560 m/s
4568 m/s ---> 4570 m/s

PRACTICE--Round off the following to three significant figures:

a.	19.87 m/s	______	d.	0.73885 m/s	______
b.	19.97 m/s	______	e.	5.001 m/s	______
c.	0.452321 m/s	______	f.	5.007 m/s	______

2. Rounding rule #2: If the calculated value falls exactly half way between to significant digits, round off to the even one. For example, 4.55 m/s, rounded off to two significant figures, would be 4.6 m/s rather than 4.5 m/sec, since 6 is even.

PRACTICE--Round off the following to three significant figures:

a.	78.85 m/s	______	d.	0.8755 m/s	______
b.	12.95 m/s	______	e.	6.005 m/s	______
c.	0.2655 m/s	______	f.	6.015 m/s	______

Now, for the 5th BIG question--Why, when insignificant digits are exactly between two digits, is the value always rounded to the even number rather than always being rounded up, down, or to the odd number?

YOUR ANSWER TO BIG Q #5:

3. Rounding rule #3: Use scientific notation to express an answer in the proper number of significant figures when necessary. For example, if the answer on your calculator turns out to be 5000000 m/s and you would like to express the answer in three significant digits, you would write the number as 5.00×10^6 m/s (can you think of another way?-there is one). Likewise, you may have to add zeros to a number written in scientific notation to communicate the proper number of significant digits. For example, if 6×10^8 was displayed on your calculator after doing a calculation, and you knew the answer should have three significant digits, you would report is as 6.00×10^8.

ACTICE--Assume the following values are displayed on your calculator after completing a calculation. How would you write each to communicate three significant figures?

a.	3000000 ______		d.	0.00003	______
b.	4×10^{23} ______		e.	3.2×10^{-23}	______
c.	40000925	______	f.	5000	______

B. PRACTICE--ASSUME THE FOLLOWING VALUES ARE DISPLAYED ON YOUR CALCULATOR AFTER MAKING A CALCULATION INVOLVING MEASUREMENTS. USE ALL OF THE ABOVE RULES TO WRITE THE VALUES IN THREE SIGNIFICANT FIGURES.

a.	56.784587	______	g.	700900	______
b.	0.00064523	______	h.	2.1×10^{-23}	______
c.	9×10^{23}	______	i.	0.000008	______
d.	4.9782×10^{23}	______	j.	0.05555	______
e.	4.155×10^{23}	______	k.	3.267×10^{-23}	______
f.	6000000	______	l.	7.995	______

C. MAINTAINING THE PROPER NUMBER OF SIGNIFICANT FIGURES IN ADDITION AND SUBTRACTION.

Suppose, in looking for Freddie's Fresh Fruit Stand, you have become lost. You stop at a convenience store and ask the clerk w to find it. The clerk tells you that if you continue east on the road in front of the store you will see a bowling alley in one mile. he fruit stand is 50 yards past the bowling alley on the left. You thank the person who helped you, but as you are walking out of the convenience store someone else in the same predicament asks you the distance to Freddie's Fresh Fruit Stand. Why would saying "It's a mile and 50 yards" be ridiculous?

Answer:

In adding and subtracting measured values the answer cannot be expressed with any more certainty than the measurement in the calculation made with the least certainty. In other words, the answer of an addition or subtraction can have no more digits to the right of the decimal point than are contained in the measurement with the least number of digits to the right of the decimal point. When solving addition and subtraction problems, add or subtract the numbers first, then round off to the decimal place of the measurement with the least number of significant figures to the right of the decimal point. For example, 3.162 m + 2.6 m is 5.762 m. However, the answer should be rounded off to 5.8 m, since the least precise measurement, 2.6 m, has only one significant digit past the decimal point.

PRACTICE: Add or/and subtract the following, and express the answer in the proper number of significant figures.

a.	3.14 m + 4.2 m	______	f.	56.798 m - 6.798 m	______
b.	3.78 + 0.075 m	______	g.	71.93 m - 40 m	______
c.	450 m + 7 m	______	h.	500.14 m + .061 m	______

***d. 3.147×10^{3} m + 4.0×10^{5} m** ______

***e. 2.147×10^{-3} m + 4.0×10^{-6} m** ______

***In adding and subtracting numbers in scientific notation, it is important to first write all values to be added or subtracted with the same exponent of ten. In this manner they can be directly compared to determine which measurement is the least precise.**

D. MAINTAINING THE PROPER NUMBER OF SIGNIFICANT FIGURES IN MULTIPLICATION AND DIVISION

The rule for maintaining the proper number of significant digits in multiplication and division is that the answer can have no more significant digits than the measurement with the fewest number of significant digits. For example, the answer to 3.454 m x 5.1 m, properly written, would have 2 significant digits (18 m^2), since 5.1 only has 2 significant digits.

PRACTICE--Solve the following, and express the answer with the proper number of significant digits. Include units in the answer!

a. 3.4 m x 58.7 m ____________

b. 245 m/520 s ____________

c. 3.45 m x 561 m x 21 m ____________

d. 8.147×10^3 m x 2.0×10^5 m ____________

e. 2.000×10^{-3} m / 4.01×10^{-6} m ____________

f. 85 m^2/20 m ____________

g. 5.6 m/.234 s ____________

h. 5 m x 61 m ____________

Congratulations! Having survived this worksheet you should now understand significant figures.