

Unit 0

REVIEW OF CHEMISTRY CONCEPTS

TOPIC 1: ATOM BASICS

TOPIC 2: SIGNIFICANT FIGURES

TOPIC 3: READING AND UNDERSTANDING SCIENCE

GLASSWARE

TOPIC 4: THE BOHR MODEL & ATOM STABILITY

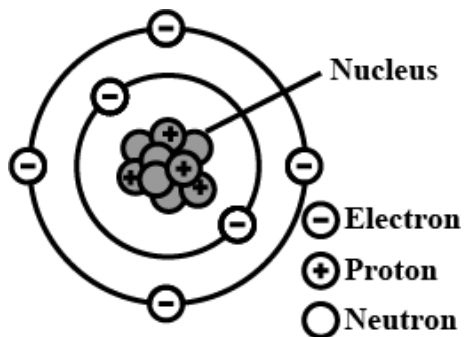
TOPIC 5: DETERMINING POLARITY

TOPIC 6: COMMON MOLECULES

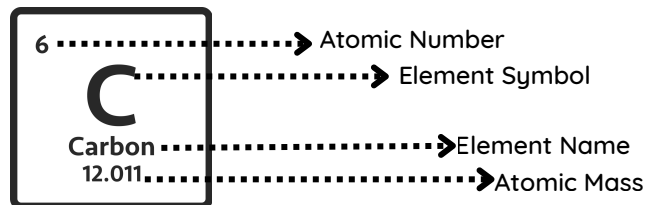
TOPIC 7: BALANCING CHEMICAL EQUATIONS

Unit 0: Review of Chemistry Concepts

Topic 1: Atom Basics



Structure of atom



Nuclear Symbol of an Element:

Mass Number = # of Protons + # of Neutrons in one atom of an element. Mass Number — A

Atomic Number = # of Protons (also equal to # of electrons in a neutral atom) Atomic Number — Z



How to calculate the number of subatomic particles:

Atomic Number =

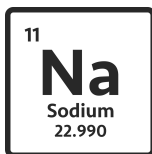
Protons =

Electrons

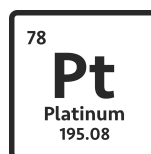
Mass Number -

Atomic Number =

Neutrons

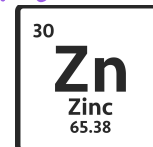


P= _____
N= _____
E= _____



P= _____
N= _____
E= _____

Practice



P= _____
N= _____
E= _____

IONS and ISOTOPES

Ions are elements that have a charge due to gaining or losing electrons.

- Cations: Positively ("paw"sitively) charged ion from losing electrons.
- Anions: Negatively charged ion from gaining electrons.

Isotopes are elements that have a different atomic mass due to gaining or losing neutrons. This does not change the type of element or give it a charge.

Isotopes



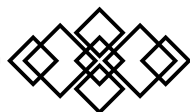
Protons: 3
Neutrons: 3



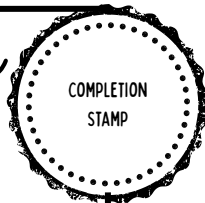
Protons: 3
Neutrons: 4



Protons: 3
Neutrons: 5



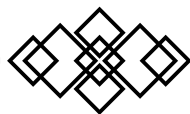
Unit 0: Review of Chemistry Concepts



Topic 1: Atom Basics Practice Worksheet

Directions: Use the information below to complete the table. Round your atomic mass values to the nearest whole number, then write the nuclear symbol for each. The first one is done for you.

Nuclear Symbol	Atom or Ion Name	Atom or Ion Symbol	# of Protons	# of Neutrons	# of Electrons	Mass Number	Atomic Number
$^{35}_{17}\text{Cl}$	Chlorine	Cl	17	18	17	35	17
		Al			13	27	
	Calcium			20	20		
	Phosphide Ion	P^{3-}		16		31	
		Rh	45	58			
	Nickel				28	59	
	Copper Ion (Cupric)		29	34	27		
		Hg				201	80
	Tin		50	69			
			36			84	
	Oxide Ion	O^{2-}	8	8			
	Barium		56				
				110		184	
				7	10		7



Unit 0: Review of Chemistry Concepts

Topic 2: Significant Figures

SIGNIFICANT FIGURES

Definition: each of the digits of a number that are used to express it to the required degree of accuracy.

- For Example: A scale that measures to one decimal place (ex: 9.2 g) cannot be read as 9.1889 because it is limited to the number of sig figs it can determine.

Rules for determining sig figs

1. All numbers 1-9 are significant. These are referred to as non-zero digits.
2. **Leading zeros** are never significant. Ex.: 0.000125 only has 3 sig figs
3. **Captive zeros** are always significant. Ex.: 2,105 has 4 sig figs
4. **Trailing zeros** are only significant if there is a decimal present.

Ex.: 12,500 has 3 sig figs 12.500 has 5 sig figs 0.0200 has 3 sig figs

$$0.750 = \underline{\hspace{2cm}}$$

$$658.0 = \underline{\hspace{2cm}}$$

$$1000.01 = \underline{\hspace{2cm}}$$

$$1,250,000 = \underline{\hspace{2cm}}$$

$$0.0080050 = \underline{\hspace{2cm}}$$

$$0.010 = \underline{\hspace{2cm}}$$

Rules for calculating with significant figures

When **multiplying or dividing**, the answer has the same number of sig figs as the value with the fewest sig figs.

Ex.: $1,205 \times 1.6 = 1,928$ Fewest sig figs = 2 Round answer to 2 sig figs = 1,900

When **adding or subtracting** the decimal places should be rounded to the same number of decimal places as the measurement with the fewest number of decimal places.

Ex.: $22.457 + 1.23 + 2.5671 = 26.2541$ Number with fewest decimal places = 1.23
Round answer to 2 decimal places = 26.25

$$\begin{array}{ccc} 32.567 & + & 135.0 & + & 1.4567 & = & \underline{\hspace{2cm}} \\ \text{L} & & \text{L} & & \text{L} & & \end{array}$$

$$\begin{array}{cc} 45.76 & * & 0.010 & = & \underline{\hspace{2cm}} \\ \text{m} & & \text{m} & & \end{array}$$

$$\begin{array}{cc} 1.678 & / & 0.42 & = & \underline{\hspace{2cm}} \\ \text{g} & & \text{mL} & & \end{array}$$

$$\begin{array}{cc} 25.771 & / & 0.012 & = & \underline{\hspace{2cm}} \\ \text{m}^2 & & \text{m} & & \end{array}$$

$$\begin{array}{ccc} 9.088 & + & 456.2 & + & 13.99 & = & \underline{\hspace{2cm}} \\ \text{cm} & & \text{cm} & & \text{cm} & & \end{array}$$

Unit 0: Review of Chemistry Concepts

Topic 2: Significant Figures Practice

COMPLETION
STAMP

Part 1: Identify the number of significant figures in each value. If applicable, identify the type of zero(s) present in the value: leading, captive, or trailing.

1.45.67: _____

2.0.00450: _____

3.1,020: _____

4.7.080: _____

5.300.0: _____

6.0.00089: _____

7.5.00 $\times 10^4$: _____

8.10,500: _____

9.0.07005: _____

10.2.300 $\times 10^{-3}$: _____

Part 2: Calculate the answer and report it with the correct number of significant figures.

11. $4.56 \times 2.1 =$ _____

12. $15.22 / 3.4 =$ _____

13. $0.00450 \times 12.7 =$ _____

14. $8.25 \times 0.300 =$ _____

15. $1.678 / 0.42 =$ _____

16. $246.24 + 238.278 + 98.3 =$ _____

17. $9.088 + 456.2 + 13.99 =$ _____

18. $45.2 - 3.89 =$ _____

19. $25.00 - 0.876 =$ _____

20. $15.67 + 2.3 =$ _____

Part 3: Round each value to the specified number of significant figures.

21. 45.6789 to 3 sig figs: _____

22. 0.0075432 to 2 sig figs: _____

23. 120.567 to 4 sig figs: _____

24. 9,876.54 to 3 sig figs: _____

25. 305.72 to 3 sig figs: _____

26. 0.00045678 to 3 sig figs: _____

27. 78.006 to 4 sig figs: _____

28. 0.090807 to 2 sig figs: _____

29. 6.022×10^{23} to 3 sig figs: _____

30. 1.999 to 2 sig figs: _____

31. 167.789 to 3 sig figs: _____

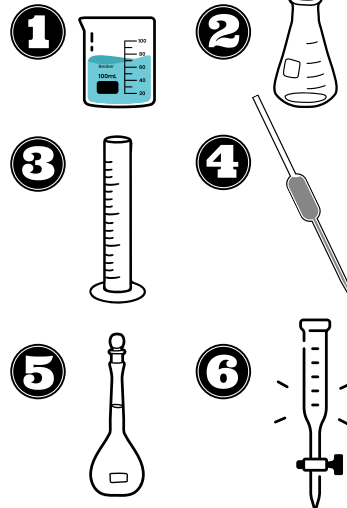
32. 25.701 to 2 sig figs: _____

33. 8562.043 to 2 sig figs: _____

Unit 0: Review of Chemistry Concepts

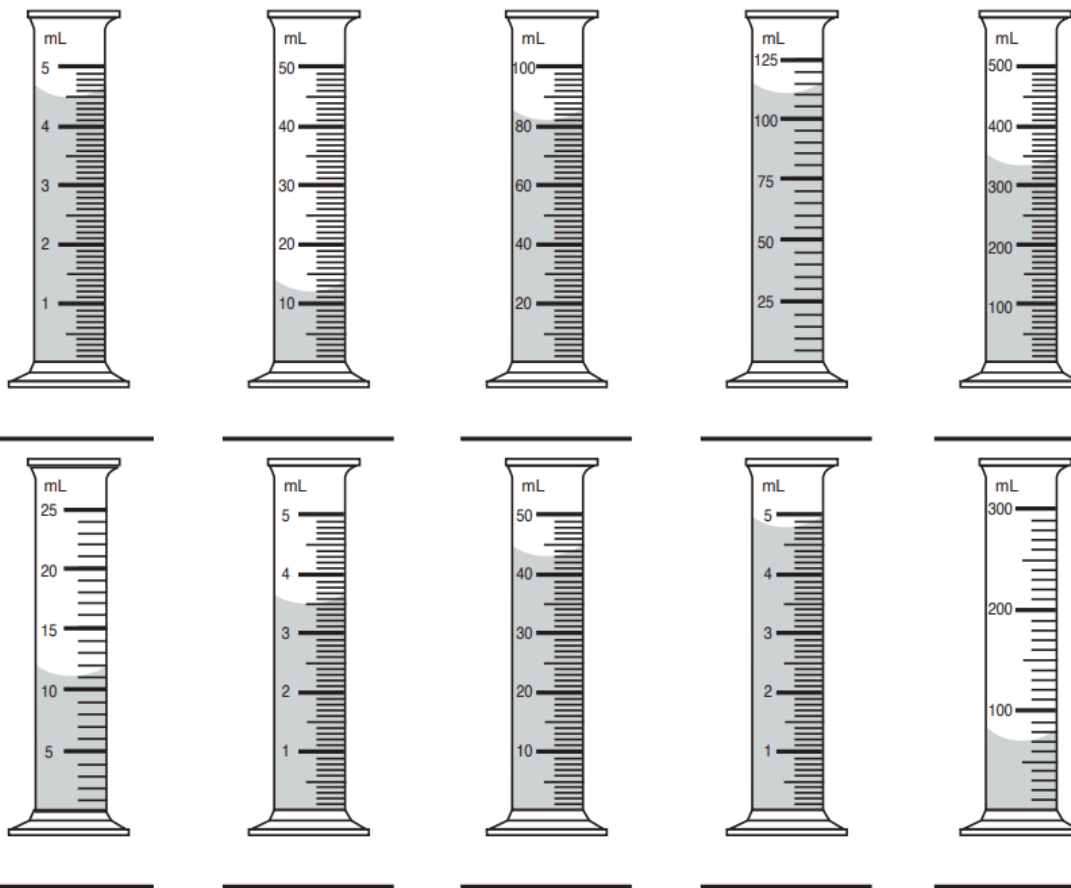
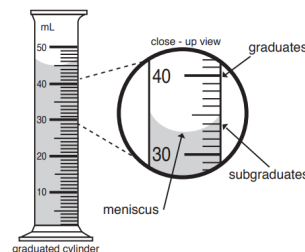
Topic 3: Reading and Understanding Science Glassware

1. **Beaker:** Can be glass or plastic; common sizes are 50 mL, 100 mL, 250 mL, and 400 mL; glass beakers may be heated. Not to be used to measure out specific volumes of solution.
2. **Erlenmeyer flask:** glass; common sizes are 100 mL, 250 mL, and 500 mL; may be heated and used in titrations.
3. **Graduated Cylinder:** glass or plastic; common sizes are 10 mL, 50 mL, and 100 mL; used to measure approximate volumes; must not be heated.
4. **Volumetric Pipette:** usually glass; it allows only for a very specific (or fixed) amount of solution to be transferred
5. **Volumetric Flask:** glass; common sizes are 25 mL, 50 mL, 100 mL, and 250 mL; carefully calibrated to a specific volume and marked with a graduation.
6. **Buret:** glass; common sizes are 25 mL and 50 mL; used to measure volumes of solutions in titrations



Reading glassware:

When reading glassware or thermometer make sure you read from eye level. Read from the bottom of the meniscus (the curve of the liquid). The degree of accuracy is unique to each piece of graduated glassware. You may read the direct measurement using the markings, and then estimate one additional significant figure. The estimate is typically a 0 or 5.



Unit 0: Review of Chemistry Concepts

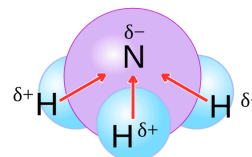
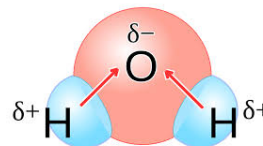
Topic 5: Determining Polarity

Polar and nonpolar are terms used to describe the distribution of electrical charge within molecules.



Polar Substances

- Contain an uneven distribution of electrons, resulting in regions of partial positive and partial negative charges.
- Caused by the atoms in a molecule having significantly different electronegativities
- The shared electrons are pulled closer to one atom, creating a dipole moment.

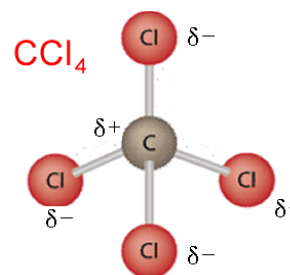


Electronegativity

Electronegativity																		8A			
1A																			8A		
H 2.1																			He		
Li 1.0	Be 1.5															B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2															Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr				
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.8	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe				
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.1	Rn				

Nonpolar Substances

- Contain a symmetrical distribution of electrons, with no significant difference in electronegativity between the atoms involved.
- Nonpolar molecules do not have permanent dipole moments and do not exhibit charged regions



Polarity and Substance Behavior

- The polarity of a substance dictates its interactions with other substances.
- For example, polar substances dissolving in polar solvents and nonpolar substances dissolving in nonpolar solvents (like dissolves like)
 - This is one reason why oil (nonpolar) and water (polar) don't mix!



Unit 0: Review of Chemistry Concepts

Topic 6: Common Molecules

1. Common Compounds: be able to recognize these!
2. Diatomic Elements: Elements that always exist in pairs (**H**ave **N**o **F**ear **O**f **I**ce **C**old **B**risk)
3. Polyatomic Ions: Ions made up of multiple elements and contain an overall charge
4. Prefixes: Used to tell you the subscript of elements within a covalent molecule

Formula	Name	Formula	Name
H ₂ O	Water	NO ₃ ⁻¹	Nitrate
CO ₂	Carbon dioxide	NO ₂ ⁻¹	Nitrite
CH ₄	Methane, carbon tetrachloride	SO ₄ ⁻²	Sulfate
CH ₃ COOH	Acetic acid	SO ₃ ⁻²	Sulfite
HCl	Hydrochloric acid	PO ₄ ⁻³	Phosphate
NH ₃	Ammonia	ClO ⁻¹	Hypochlorite
H ₂	Hydrogen gas	ClO ₂ ⁻¹	Chlorite
N ₂	Nitrogen gas	ClO ₃ ⁻¹	Chlorate
F ₂	Fluorine gas	ClO ₄ ⁻¹	Perchlorate
O ₂	Oxygen gas	CH ₃ COO ⁻¹	Acetate
I ₂	Iodine gas	C ₂ H ₃ O ₂ ⁻¹	Acetate
Cl ₂	Chlorine gas	CO ₃ ⁻²	Carbonate
Br ₂	Bromine gas	CrO ₄ ⁻²	Chromate
OH ⁻¹	Hydroxide (base)	MnO ₄ ⁻	Permanganate
H ₃ O ⁺	Hydronium (acid)	NH ₄ ⁺¹	Ammonium

X₁ or X = Mono- X₂ = Di- X₃ = Tri-
 X₄ = Tetra - X₅ = Penta- X₆ = Hexa-
 X₇ = Hepta- X₈ = Octa- X₉ = Nona-



Unit 0: Review of Chemistry Concepts

Topic 7: Balancing Chemical Equations

According to the law of conservation of mass, matter cannot be created or destroyed. An equation must represent equal amounts of atoms of the same element on both sides of the reaction. For example, if you have 4 carbon atoms in the reactants, you must have 4 carbon atoms in the products.

Since each chemical reaction is unique there are no set steps on how to balance a chemical equation.

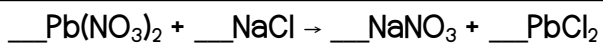
A couple of things to note when balancing:

1. You can only change the coefficient. You CANNOT change the subscript of a compound.
2. Polyatomic ions should be counted as one "group" if they remain unchanged.
3. Water can be rewritten as H-OH to help in balancing.
4. Everything after the coefficient is multiplied by that value. If there is a blank in front of a compound then the coefficient is 1, NOT a 0.
5. Start by counting the initial number of elements on each side. Change coefficients as needed, recount, and repeat until you have the same number of elements on each side.

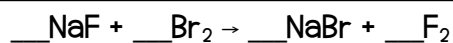
LABEL THE KEY PARTS OF A
CHEMICAL REACTION:



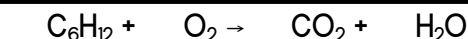
EXAMPLE:



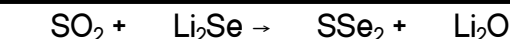
Type of rxn: _____



Type of rxn: _____



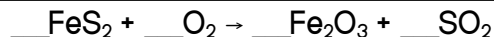
Type of rxn: _____



Type of rxn: _____



Type of rxn: _____



Type of rxn: _____

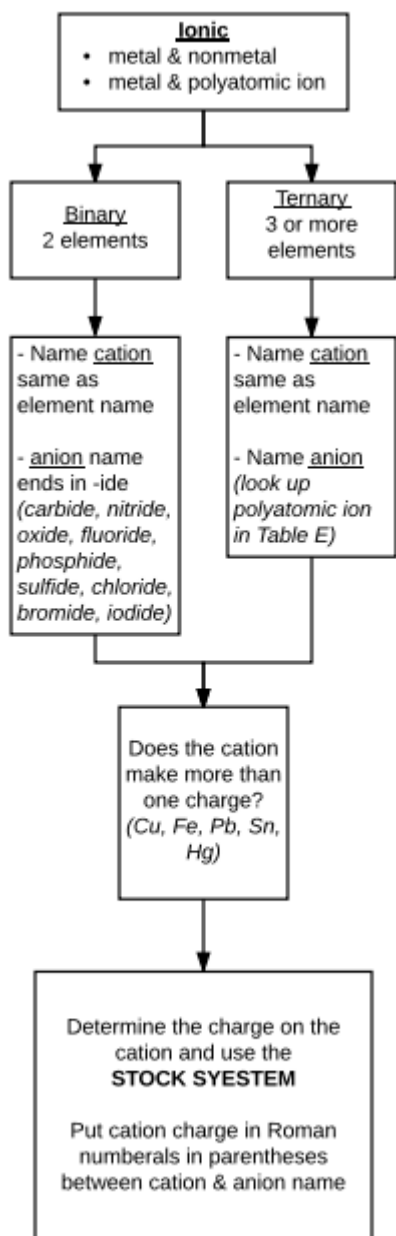
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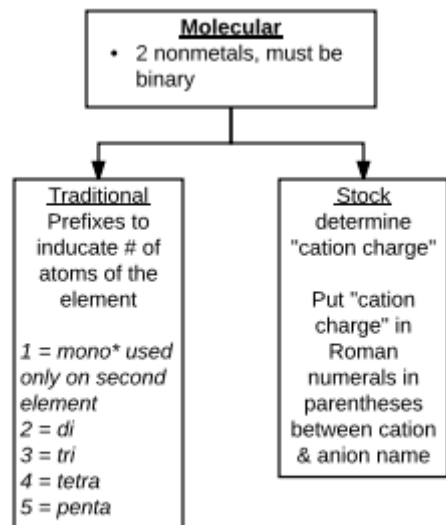
FORMULA WRITING

- ✓ Total charge of a compound is zero.
- ✓ Use oxidation numbers/states (charges) on the top right hand corner of the Periodic Table.
- ✓ Table S for element names and symbols! Table E for Polyatomic Ions!
- ✓ (s) = solid, (l) = liquid, (g) = gas, (aq) = aqueous (solution, dissolved in water)

Naming Ionic Compounds



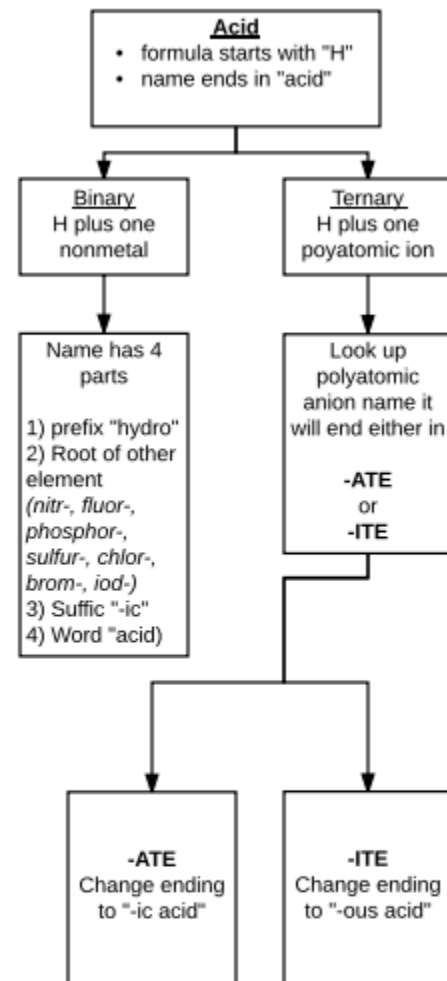
Naming Covalent Compounds



Writing Formulas

- Write the symbols
- If molecular, use prefixes to determine how many of each element you have
- If ionic, look up the charges
 - Criss-Cross and Drop the Sign (Use parentheses for more than one polyatomic ion)
 - If a roman number is used, that is the charge of the metal

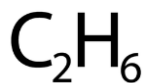
Naming Acids



*Add the extra syllable to sulfur and phosphorous containing ions

Types of Formulas

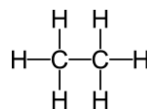
Molecular



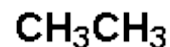
Empirical



Structural



Condensed



Write formulas for the following compounds:

lithium chloride _____

ammonium permanganate _____

silver nitrate _____

zinc hydroxide _____

carbon disulfide _____

iron(III) phosphate _____

copper(I) iodide _____

tin(IV) fluoride _____

barium dichromate _____

beryllium nitrite _____

Write names for the following compounds:

NaClO_4 _____

P_2O_3 _____

Li_3P _____

KHSO_4 _____

FeS _____

PbCrO_4 _____

MgBr_2 _____

ZnSO_4 _____

K_2CO_3 _____

Cl_2S_5 _____

Al_2O_3 _____

Use dimensional analysis (the “factor-label” method) to solve the following problems. **Show all steps** needed to convert from starting units to ending units. **Indicate all relationships needed** before setting up and solving the problem. Use any of the following relationships if needed:

1 mile = 1760 yds	16 oz = 1 lb	1 L = 1.057 qts	1 day = 24 hours
1 yd = 3 ft	2000 lbs = 1 ton	4 qts = 1 gal	1 hour = 60 mins
1 ft = 12 in	1 oz = 28.35 g	32 oz = 1 qt	1 min = 60 secs
1 mile = 1.6093 km	1 kg = 2.205 lbs	1 qt = 2 pts	

1. A runner competed in a 5-mile run. How many yards did she run?
2. In the Tour de France, cyclists ride 3,653.6 km in 20 days. How many miles do they go? [Hint: watch for unimportant information!]
3. After a nice meal, perhaps you’d finish it off with a pound (1.00 lb) cake for dessert. What would the name of this cake be in grams?
4. In the US milk is sold by the gallon, while in Denmark it is sold by the liter. How many liters of milk would you need to equal one gallon?
5. If you go to school for 180 days each school year and each school day is 7 hours long, how many hours are spent in school in one school year?