REVIEW of Grade 11 Chemistry

SCH4U_08 - 09 NAME: ________________

Section A: Review of Rules for Significant Digits

All measurements have a certain degree of _________________ associated with them.
All the accurately known digits and the one uncertain digit are called ____________ _____________ or ____________ _____________.
The last figure shows where uncertainty begins.
If a thermometer indicates a boiling point of 36.2 ^{0}C, and has an uncertainty of ± 0.2 ^{0}C, all figures in that 36.2 are significant, including the 0.2, which is uncertain.
The notation 36.2 ± 0.2 ^{0}C indicates how uncertain it is.

1. All measured nonzero digits are significant, e.g. 123, 12, 123.3

2. Zeros in between non-zero digits are significant figures, e.g. 102, 12.0, 102.3

3. Zeros appearing in front of all nonzero digits are not significant. They are acting as placeholder, e.g. 0.123, 0.0123, 0.00123

4. Zeros at the end of a number and to the right of a decimal point are significant, e.g. 123.0, 1.10, 1.0

5. Zeroes at the end of a number without a decimal point are ambiguous. Adding a decimal point indicates their significance, e.g. 1230. 120. 20

6. When **adding or subtracting** numbers, **the final answer should have the same number of decimal places** as the measurement having ____, (i.e. the measurement with the least precision).

7. When **multiplying and dividing** numbers, the **final answer should have the same number of significant digits** as the measurement having the _______________________________, (i.e. the measurement with the least accuracy).

8. When there are a series of calculations to do to obtain the final result, ____.
(Or, if you do round off, leave at least one extra digit until the end of all the calculations.)

Now answer the following questions:

1. How many significant digits are there in each of the following measurements?

a) 204.45 ha ____________ b) 18.23 s _____________

c) 380 000 km ____________ d) 0.00560 g _____________

2. Express the answer to each of the following calculations with the **correct number of significant digits and using proper scientific notation.** *(1 mark each)*

a) 13.89cm + 6.7732 cm _____________ b) 120 km^3 / 8.56 km _____________

c) 3.0899 mm^2 x 22.4 mm _____________ c) 3.3 x 10^{-6} m x 1.05 x 10^2 m _____________

Section B: Review of Nomenclature

1. What System to use?

	System				**Example**
First	**IUPAC**		**Classic**		
Element	**STOCK**	**None**	**Prefix**	**– OUS / – IC**	
Metal: only 1 oxidation #		✔			Na^{+1}, K^{+1}, Ca^{+2}
Metal: more than 1 oxidation # (but only max. 2 possible)	✔		✔	✔	Cu^{+1}/ Cu^{+2}, Pb^{+2} / Pb^{+4}, (not: Mn, Cr, V, etc.)
Metal: more than 1 oxidation # (but more than 2 possible)	✔		✔		Mn^{+2}, Mn^{+4}, Mn^{+6}, Mn^{+7}
Non-metal	✔		✔		NO, N_2O, N_2O_4

2. Binary Compound: use the suffix: " ____________".
   (Exceptions: hydroxide, OH^{-1}, and cyanide, CN^{-1})

e.g. CaO, $ZnCl_2$, Fe_2O_3, BN, $Ba(OH)_2$, KCN

3. Polyatomic ions: " — ate", " — ite"
   E.g. Na_2CO_3, $FeSO_4$

4. Hydrate: name the compound + use Greek numbering to indicate the molecules of water + hydrate.
   E.g. $Na_2CO_3 . 10\ H_2O$, $FeSO_4 . 6\ H_2O$

5. Binary acids: "hydro ________________ --- ic acid".
   E.g. $HCl_{(aq)}$, $HF_{(aq)}$, $H_2S_{(aq)}$

6. Polyatomic acids: $[H_n X O_n]$:

	—ic acid,	—ous acid,	hypo_____ous acid,	Per____ic acid
E.g.	$HClO_{3\ (aq)}$,	$HClO_{2\ (aq)}$,	$HClO_{(aq)}$,	$HClO_{4\ (aq)}$,

7. Ion form the polyatomic acid:
   — ate, —ite, Hypo____ite , Per _____ate

Now answer the following questions:

1. Name the following compounds.

a) Mg_3N_2 ________________ b) $AgCN$ ________________

c) $Ca(NO_3)_2$ ______ ____________ d) $H_2SO_{4\,(aq)}$ ________________

e) $HCl_{(aq)}$ ________________ f) $CuSO_4.5H_2O$ ________________

2. Give the formula for each compound.

a) cupric nitrate dihydrate ____________ b) dintrogen trisulphide ____________

c) magnesium carbide ____________ d) ammonium phosphate ____________

e) chromium (III) bromide ____________ f) plumbous acetate ____________

Section C: Review of Balancing Equations and Types of Reactions

Balance the following equations. Identify the **type of reaction** occurring, (i.e. synthesis, decomposition, single displacement, double displacement).

1. __ $Al_{(s)}$ + __ $NiCl_{2\,(aq)}$ $\longrightarrow$ __ $AlCl_{3\,(aq)}$ + __ $Ni_{(s)}$

2. __ $Mg(OH)_{2\,(aq)}$ + __ $HNO_{3\,(aq)}$ $\longrightarrow$ __ $Mg(NO_3)_{2\,(aq)}$ + __ $H_2O_{(l)}$

3. __ C_3H_8 + — O_2 $\longrightarrow$ — CO_2 + — H_2O

4. Write a **balanced equation** for the complete combustion of octane, $C_8H_{18\,(g)}$ in the presence of excess oxygen gas, include the state symbols.

5. When an aqueous solution of copper (II) sulphate is added to an aqueous solution of sodium hydroxide, a blue-green precipitate is obtained.

6. When an aqueous solution of calcium chloride is added to an aqueous solution of potassium phosphate, a white precipitate is obtained.
7. A solution of lead (II) nitrate is mixed with a solution of sodium chromate. A precipitate forms in the presence of a soluble solution.

Section D: Review of Solubility Rules and Net- Ionic Equations

Solubility Rules:

Steps for writing net-ionic equation:

Now for **questions 5, 6 and 7 from Section C** above, write the**:**

(a) name and the **formula** for the possible **precipitate,**

(b) balanced chemical equation for the reaction described, include **state symbols.**

(c) balanced total dissociated ionic equation,

(d) balanced net-ionic equation.

Section E: Review of Chemical Calculations

The mole is the unit in which amounts of substance are measured in chemistry.

The mole is defined as that amount **of substance that contains the same number of particles as there are atoms in exactly 12 g of the isotope carbon 12.**

The number of particles in a mole is found to be $6.02 * 10^{23}$ this number is called the Avogadro constant and has the symbol N_A

n	number of moles
m	mass (g)
m	Molar mass (g mol^{-1})
N	number of entities
N_A	Avogadro's number
C	molar concentration (mol L^{-1}, mol dm^{-3})
V	volume (L, dm^3)
V_M	molar volume of gas (22.4 L at STP)

The relationships between amount of substance, number of particles, mass of solid, and volume of gas are very important:

amount		**number of particles**		**mass of solid**		**volume of gas**
1 mole	=	$6.02 * 10^{23}$	=	A_R or M_R in grams	=	22.4 dm^3 at STP

Many calculations involve converting from one part of this relationship to another; always go back to this key line at the start of your calculation.

When performing chemical calculations remember the following:
(i) To define the particles you are talking about.
E.g. Is your mole of oxygen $6.02 * 10^{23}$ oxygen **atoms** which weigh 16 g or $6.02 * 10^{23}$ oxygen **molecules** which weigh 32 g?
(ii) Substances are often not pure, but are diluted in solutions.
The quantity of substance in a solution is called its concentration, expressed in several different ways:
- grams per liter shortened to g/L or g L^{-1}
- grams per cubic decimeter short to g/dm^3 or g dm^{-3}
- moles per liter shortened to mol/L or mol L^{-1}
- moles per cubic decimeter shortened to mol/dm^{-3} or mol dm^{-3}

(iii) **Volumes are measured in several different units**: 1 dm^3 = 1 liter = 1 000 cm^3 = 1000 mL

Steps for Calculating Empirical and Correct Molecular Formula

Steps to Stoichiometry:

Now answer the following questions:

1. i) How many moles in 5.00×10^2 g of iron?

   ii) How many iron atoms in 5.00×10^2 g of iron?

   iii) How many moles are there in 185 g of calcium hydroxide ?

   iv) How many molecules are there in 196 g sulphuric acid?

   v) How many oxygen atoms, sulphur atoms, hydrogen atoms and sulphate ions, hydrogen ions are contained in 196 g sulphuric acid ?

2. Calculate the mass of:

   i) 1.50 moles of oxygen gas, O_2

   ii) 750 cm^3 of 0.0150 mol dm^{-3} NaOH

   iii) 7.00 mol potassium fluoride, KF

   iv) 3.01×10^{22} molecules of nitric acid, $HNO_{3\ (aq)}$

   v) 5.62 L of carbon dioxide gas at 65.5 ^{O}C and 126 kPa.

3. The molar mass of a compound with the empirical (simplest) formula CH_2O was found to be 240 g mol^{-1}. What is the molecular formula of the compound ?

4. The percentage composition of tartaric acid is: 32.01 % C, 4.03% H, and 63.96 % O. Given that the molecular mass of tartaric acid is 150 amu, determine its **molecular formula.**

5. Using the equation below, how many grams of ammonia will be formed if 75.0 g of nitrogen reacts with excess hydrogen? $N_{2\ (g)} + H_{2\ (g)} \longrightarrow NH_3$ **(BALANCE)**

6. A mixture of 5.00 g of $H_{2\,(g)}$ and 10.0 g of $O_{2\,(g)}$ is ignited. Water forms according to the following equation: $2H_{2\,(g)} + O_{2\,(g)} \longrightarrow 2H_2O_{(l)}$

a) Which reactant is limiting?

b) How much water will be produced by the reaction?

7. Calcium carbonate 'fur' on the inside of a kettle used in a hard water area of the country can be removed using a dilute solution of hydrochloric acid. What volume of 0.010 mol l^{-1} hydrochloric acid would be needed to remove 2.00 g of calcium carbonate from the kettle? The equation for the reaction is:

$$CaCO_{3(s)} + 2\,HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

8. One of the reactions involved in the smelting of copper sulphide ores involves copper (I) oxide and copper (I) sulphide involves the following equation:

$$2\,Cu_2O_{(s)} + Cu_2S_{(s)} \longrightarrow 6\,Cu_{(s)} + SO_{2\,(g)}$$

If 50.0 g of Cu_2O is heated with 25.8 g of Cu_2S , then:
a) determine which reagent, if any, is in excess.
b) calculate the theoretical yield of copper.
c) determine the percent yield if 58.0 g of copper is actually obtained.

Section F: Review of Atomic and Ionic Structure

1. Write the **complete,** and **short-hand** electronic configuration (in terms of s, p, d) for:
(i) $_{15}P$ ____________________________ (ii) $_{26}Fe$ ____________________________

2. Hydrogen has three isotopes: $_1H$, $_1H$, $_1H$. Give the number of **protons, neutrons** and **electrons** found in each isotope.

3. Which of the following elements will have the **largest** atomic radius?
a) Cs, K, or Li __________ b) F, B or Li __________
c) K^{+1}, Mg^{+1}, Al^{+3} __________ d) O, O^{-1}, O^{-2} __________

4. Which of the following will have the **smallest** first ionization potential energy?
a) Li, B, F ______________ b) Si, S, Sb ______________

The difference in electronegativity, ΔEn, can be used to describe the chemical bond as:
i) ionic, ΔEn > 1.7
ii) polar covalent, ΔEn : 0.5 – 1.7
Iii) non-polar covalent, ΔEn < 0.5

In a polar molecule the bond dipoles do NOT cancel, e.g. H_2O

5. Given the following combinations of elements and their electronegativities, state **what kind of bond** (ionic, polar covalent, or covalent) is formed.
a) potassium (0.9) and chlorine (2.9) ____________________
b) hydrogen (2.1) and oxygen (3.5) ____________________
c) two sulphur atoms (2.4) ____________________
d) phosphorus (2.l) and chlorine (3.0) ____________________

6. Explain the difference between **electron affinity** and **electronegativity**, give an **example** in each case.

7. Draw the Lewis structure for each of the following molecules and state the **shape** and indicate if the molecule is **polar or non-polar**.
a. Chloroform, $CHCl_3$

b. Ammonia, NH_3

c. Water, H_2S

8. Of the chemical substance listed below:
$CH_{4\ (g)}$ $MgCl_{2\ (aq)}$ $CCl_{4\ (g)}$ $HI_{\ (aq)}$ $KOH_{(aq)}$

State which is:
a base _________
an acid _________
an organic compound: _____________
which will be good conductors of electricity: _______________

9. Explain what is meant by a strong electrolyte and a weak electrolyte. Give an example of each.

10. Explain the following:
a. Give **two** reasons why the lithium ion, Li^+, has a smaller radius than the lithium atom.

b. Explain which ion is smaller: Na^{+1} or F^{-1}

c. Give **two** reasons why noble gases are not assigned electronegativity values.