Name _____

Date _____ Bell _____

Chemistry Extended Learning Packet

I. Significant Figures and Scientific Notation

a. Rules

1. Count all non-zero integers
2. Zeros:
    a. Leading zeros that precede all of the nonzero digits NEVER count.
    b. Zeros in between non-zero integers ALWAYS count.
    c. Trailing zeros to the right of non-zero digits only count if there is a decimal point in the figure.

b. *Problems*: Count the number of significant figures.

1. 0.0230 _____
2. 143,000,210 _____
3. 0.00038 _____
4. Round 342,593 to 3 significant figures _____
5. Round 0.004583 to 2 significant figures _____

II. Calculating with significant digits

a. Multiply/Divide

1. The number of significant figures in the result is the same as that in the measurement with the smallest number of significant digits.
2. Example: 1.2 x 2.048 = 2.5 (2 sig figs)

b. Add/Subtract

1. The limiting term is the one with the smallest number of decimal places.
2. Example. 32.015 + 12.1 + 0.01 = 32.1

c. Problems - Calculate

1. 2.09 x 3,309 = _____
2. 12.47 =0.0145 = _____
3. 24.256 + 143 = _____
4. 8,008.1 – 234 = _____

III. Scientific Notation

a. Scientific Notation is used to express very large and very small numbers as the product of a number between 1 and 10 and a power of 10.

b. Example

1. $93{,}000{,}000 = 9.3 \times 10^{7}$
2. $0.000305 = 3.05 \times 10^{-4}$

c. Problems: Convert to scientific notation or to standard form.

1. 6,0430,000,000,000 _____
2. 0.0000302 _____
3. 6,074,720,000 (2 sig figs) _____
4. 1.4×10^{5} _____
5. 8.406×10^{-8} _____

IV. Dimensional Analysis

a. Metric Conversions Problems (Conversion factors on pg 118 in text)

1. 375 mm = _____ m
2. 0.0035 kg = _____ g
3. 1.753 mm = _____ km

b. Dimensional Analysis

1. Set up t-charts to cancel out unwanted units
2. Example: 62 cm = _____ in (1 in = 2.54 cm)

$$\frac{62.0 \text{ cm}}{} \left| \frac{1 \text{ in}}{2.54 \text{ cm}} \right. = 24.4 \text{ in}$$

c. Problems: Convert using conversion chart in Appendix F in your textbook

1. 723.5 g = _____ lb
2. 26.2 mi = _____ m
3. 2.23 m = _____ yd

d. Grams ↔ moles ↔ # atoms/molecules

1. Example: 1.45 g CO_2 = _____molecules CO_2

$$\frac{1.45\text{ g } CO_2}{} \left| \frac{1\text{ mol } CO_2}{44.01} \right| \frac{6.022 \times 10^{23}\text{ mlc } CO_2}{1\text{ mol } CO_2} = 1.98 \times 10^{22}\text{ mlc } CO_2$$

e. Problems: Calculate

1. Calculate the number of moles of water in 13.8 g of water.

2. Calculate the number of grams in 1.67×10^{24} mlc of NO_3.

3. Calculate the number of Ti atoms in 0.94 moles of Ti.

V. Balancing Chemical Equations

a. Balance the equations:

1. _____$Na_{(s)}$ + _____$HCl_{(aq)}$ → _____$H_{2(g)}$ + _____$NaCl_{(s)}$

2. _____$NH_{3(g)}$ + _____$O_{2(g)}$ → _____$NO_{(g)}$ + _____$H_2O_{(g)}$

3. _____$K_{(s)}$ + _____$H_2O_{(l)}$ → _____$H_{2(g)}$ + _____$KOH_{(aq)}$

VI. Stoichiometry

a. See your limiting reactant flip chart and the Chemical Quantities packet for examples.

b. Problems: Calculate

1. Calculate the mass of carbon dioxide produced from 12.6 g of octane, C_8H_{18}, in the following reaction.

$$2C_8H_{18(g)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(g)}$$

2. The balanced equation $P_{4(s)} + 6\ H_{2(g)} \rightarrow 4\ PH_{3(g)}$ tells us that 3 mol H_2 would
   a. React with 1 mol P_4
   b. Produce 4 mol PH_3
   c. not react with P_4
   d. Produce 2 mol PH_3
   e. React with 2 mol P_4

3. Calculate how many grams of $CaCl_2$ would be obtained if 17.5 g HCl is allowed to react with excess $CaCO_3$ in the reaction:

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