Announcements

Wednesday, September 30, 2009

Exam 1 will be graded sometime over the weekend. Scores will be entered in D2L. I'll send an email to the class when they're up.

Answer key and blank for Exam 1 are in D2L now.

MasteringChemistry due dates:

- Ch 4: Fri, Oct 9
- Ch 5: Fri, Oct 23
- Ch 6: Fri, Oct 30

Lab report due dates:

- Exp 5: Mon, Oct 12
- Exp 6: Mon, Oct 19

Chapter 4: Stoichiometry and aqueous reactions

Stoichiometry: amounts of substances in balanced chemical reactions

$\underline{2}\ C_8H_{18} + \underline{25}\ O_2 \rightarrow \underline{16}\ CO_2 + \underline{18}\ H_2O$

$\underline{2}\ C_8H_{18}$ molecules + $\underline{25}\ O_2$ molecules → $\underline{16}\ CO_2$ molecules + $\underline{18}\ H_2O$ molecules

$\underline{2}$ mol C_8H_{18} + $\underline{25}$ mol O_2 → $\underline{16}$ mol CO_2 + $\underline{18}$ mol H_2O

How many moles of CO_2 can be produced from the combustion of 8.7 mol octane (C_8H_{18})?

(Use coefficients from balanced chemical equation to make a mole ratio)

$$8.7\ \text{mol}\ \cancel{C_8H_{18}} \times \frac{16\ \text{mol}\ CO_2}{2\ \text{mol}\ \cancel{C_8H_{18}}} = 70.\ \text{mol}\ CO_2 \text{ or } 7.0 \times 10^1\ \text{mol}\ CO_2$$

How many grams of CO_2 can be produced from the combustion of 4.50×10^4 g of octane?

$$\text{g oct} \rightarrow \text{mol oct} \rightarrow \text{mol}\ CO_2 \rightarrow \text{g}\ CO_2$$

$$4.50 \times 10^4\ \text{g oct} \times \frac{1\ \text{mol oct}}{114.224\ \text{g oct}} \times \frac{16\ \text{mol}\ CO_2}{2\ \text{mol oct}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2} = \boxed{1.39 \times 10^5\ \text{g}\ CO_2}$$

Limiting reactants

Whenever amounts of more than one reactant are known, you must find the limiting reactant

- Which reactant is consumed first?
- Calculate moles of a single product that each reactant will make
- Reactant that makes fewer product moles is the limiting reactant.
- That's how many product moles can be formed.

$Zn + 2\ HCl \longrightarrow ZnCl_2 + H_2$ pdt

.30 mol   .52 mol HCl   ? mol

If 0.30 mol Zn react with 0.52 mol HCl, how many mol H_2 can be formed?

$.30 \text{ mol Zn} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Zn}} = \cancel{.30 \text{ mol } H_2}$ Cannot be produced

$.52 \text{ mol HCl} \times \frac{1 \text{ mol } H_2}{2 \text{ mol HCl}} = \boxed{.26 \text{ mol } H_2}$ can be produced

runs out first

HCl is limiting reactant

To calculate amount of leftover reactant, first calculate amount of excess reactant that was actually used, then subtract the given amount.

$.26 \text{ mol } H_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol } H_2} = .26 \text{ mol Zn}$ was used

$.30 \text{ mol Zn (start)} - .26 \text{ mol Zn (used)} = \boxed{.04 \text{ mol Zn}}$ remain unreacted

Limiting reactant with starting masses

If 7.36 g Zn react with 6.45 g S_8, how many grams ZnS can be produced? What mass of reactant remains?

$8\ Zn + S_8 \longrightarrow 8\ ZnS$

lim react.

$7.36 \text{g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{g Zn}} \times \frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} = .11257 \text{ mol ZnS}$

$6.45 \text{g } S_8 \times \frac{1 \text{ mol } S_8}{256.5 \text{g } S_8} \times \frac{8 \text{ mol ZnS}}{1 \text{ mol } S_8} = \cancel{.201 \text{ mol ZnS}}$ cannot form

$.11257 \text{ mol ZnS} \times \frac{97.45 \text{g ZnS}}{1 \text{ mol ZnS}} = \boxed{11.0 \text{ g ZnS}}$ can be produced

$.11257 \text{ mol ZnS} \times \frac{1 \text{ mol } S_8}{8 \text{ mol ZnS}} \times \frac{256.56 \text{ g } S_8}{1 \text{ mol } S_8} = 3.61 \text{ g } S_8$ used

$6.45 \text{g } S_8 - 3.61 \text{ g } S_8 = \boxed{2.84 \text{ g } S_8}$ unreacted

Theoretical yield: calculated product mass from stoichiometry calculation

(upper limit) no more than 11.0 g ZnS (above)

Actual yield: measured product mass from experiment

actual < theoretical   given in problem or measured in lab

$$\%\ \text{Yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\%$$

Say actual yield was 10.2g $\% = \frac{10.2g}{11.0g} \times 100\% = \boxed{92.7\%}$ yield

Solution concentration and stoichiometry

Most chemical reactions in this course take place in solution (dissolved in water)

Any concentration = $\frac{\text{amount solute}}{\text{amount solution}}$ total

Molarity (M) = $\frac{\text{moles of solute}}{\text{liters of solution}}$

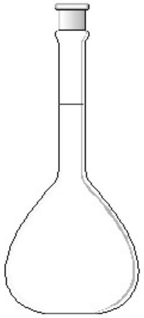
For instance, a 0.50 M (molar) $AgNO_3(aq)$ solution is called a 0.50 "molar" silver nitrate solution

1 liter of this solution contains .50 moles of $AgNO_3$

How do you prepare 250.0 mL of a 0.10 M $AgNO_3(aq)$ solution? (Use the given molarity as a conversion factor between moles solute and liters solution)

0.10 M means $\frac{.10 \text{ moles } AgNO_3}{1 \text{ L solution}}$

0.10 mol/L



volumetric flask