Chem 101 Key for Test #2 Review Guide.

Review Session is Tuesday 2/15/05, at 1:30-2:15 pm sharp in PS 607.

Midterm #2 is on Friday, Feb. 18, 2005. Test #2 will focus on Chapters 4-5
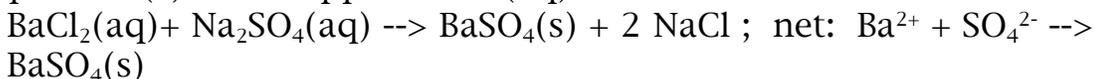
Sample short problems:

1) answer: #g Na_2CO_3 needed=
   .250L(.14molNa/L)(1molNa$_2$CO$_3$/2molNa)(106g/mol)=1.86g

2) a) sodium sulfate & barium chloride. React? _yes_ (yes/no)
   product(s): _BaSO4 ppt & NaCl(aq)_
   $BaCl_2(aq)+ Na_2SO_4(aq)$ --> $BaSO_4(s)$ + 2 NaCl ; net: $Ba^{2+} + SO_4^{2-}$ -->
   $BaSO_4(s)$
   b) silver chloride & potassium nitrate. React? _no_ (yes/no) If so,
   product(s): _None since AgCl is insoluble and will not be able to go
   into solution._
   c) Lead acetate & ammonium iodide. React? _yes_ (yes/no) If so,
   product(s): _PbI$_2$_ , _____.
       $Pb(CH_3CO_2)_2$ (aq) + 2NH$_4$I (aq) → $PbI_2(s)$ + 2 $NH_4CH_3CO_2(aq)$. Net
   ionic: Pb^{2+} + 2I$^-$ -->PbI_2
   d) sulfuric acid + sodium hydrogen carbonate ? _yes_ (yes/no) If so,
   product(s): _Na$_2$SO$_4$_ , _CO$_2$(g), H$_2$O(l)_.
       H_2SO_4 (aq) + 2 NaHCO$_3$ (aq) --> $Na_2SO_4(aq)$ + 2 $CO_2(g)$ + 2 H_2O (l)
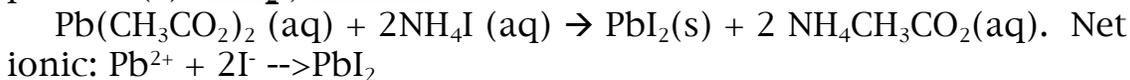
3)) Which of the following are redox reactions? Acid base reactions?
   Balance them anyway.a) Cu^{2+} (aq) + Ag(s) -> Cu(s) + Ag$^+$(aq)
   Redox? _yes_ (yes/no)
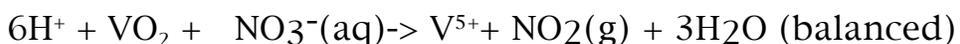 $Cu^{2+}(aq)$ + 2Ag(s) -> Cu(s) + 2Ag$^+$(aq) (balanced)
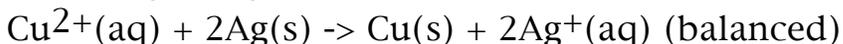
   b) metathesis reactions in general.   Redox? ___NO_ (yes/no)
      (Note: metathesis = double replacement = exchange reactions)
   c) $VO_2(s)$ + HNO$_3$(aq)-> V(NO$_3$)$_5$(aq) + NO$_2$(g) (unbalanced)
      Redox? _yes_ (yes/no)
   6H$^+$ + VO$_2$ + NO$_3^-$(aq)-> V^{5+}+ NO$_2$(g) + 3H$_2$O (balanced)

   d) CH$_4$ + O$_2$-> CO + H$_2$O (unbalanced)   Redox? _yes_ (yes/no)
   2CH$_4$ + 3O$_2$-> 2CO + 4H$_2$O(balanced)
   e) H_2CO_3 + NaCl--> Na_2CO_3 + HCl Redox? _NO_ (yes/no) (their oxidation
   numbers don't change) It's acid base since there is a transfer of
   protons from H_2CO_3 to Cl$^-$.
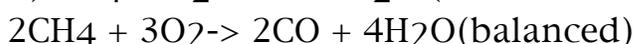
f) $H_2CO_3 \rightarrow H_2O + CO_2$? Redox? <u>NO</u> (yes/no) (their oxidation numbers don't change, and no transfer of protons)
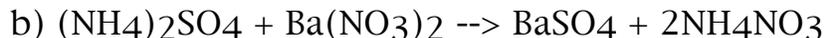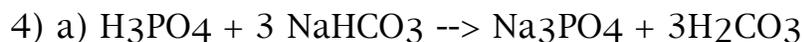
4) a) H3PO4 + 3 NaHCO3 --> Na3PO4 + 3H2CO3
b) (NH4)2SO4 + Ba(NO3)2 --> BaSO4 + 2NH4NO3
c) This is an redox reaction. How do we know?
OK, write the equation and determine the oxidation numbers (ON):

$Na_2 C_2 O_4$ + K Mn O$_4$ --> Mn^{2+} + C O$_2$

ON: +1 +3 –2   +1 +7 –2   +2   +4 –2   (note these ON's are per atom)
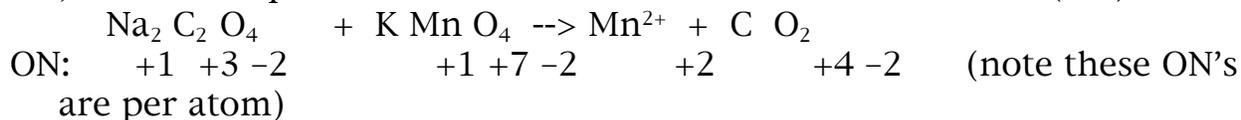
Since the ON of C and of Mn change, we can safely conclude that this is a redox reaction. To balance this, first write the 2 half reactions ("1/2 rxn"). Train yourself to do this in the <u>net ionic</u> form:

Oxidation 1/2 rxn: $C_2O_4{}^{2-}$ --> $2CO_2$ +2e^{2-} (note: I had to balance the C's first before determining how many e's to place on the right side. Remember that in oxidations, the electrons are on the right side).
We note that this half reaction equation is ALREADY BALANCED.

Reduction 1/2 rxn: $MnO_4{}^-$ + 5e---> Mn^{2+} (note that the ON for Mn went down from +7 to +2. That is a gain of 5 e-'s.   this is still unbalanced. So here we need to balance the oxygens with H_2O and after that, the H's with H$^+$. Note that this is the correct rule for "acidic conditions". If it were basic conditions, we normally add 2 H_2O's for each O to balance and the H's with OH$^-$.)

Enough instructions, let's balance the second equation: first balance the 4 O's on the left side by adding 4 H_2O's on the right side:

$MnO_4{}^-$ + 5e---> Mn^{2+} + 4H$_2$O
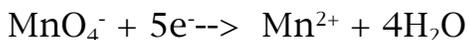
OK, now although the O's are balanced, a new problem arises, we have added 8 H's on the right side. To balance those, we add 8 H='s on the left side:

8H$^+$ + MnO$_4{}^-$ + 5e---> Mn^{2+} + 4H$_2$O (finally this half reaction is balanced!)
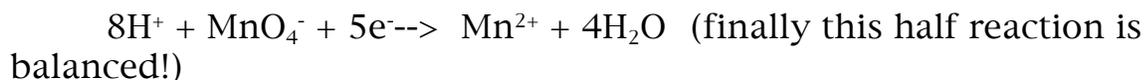
finally, check to see if the charges are balanced. Yes they are! Both sides have a net +2 charge.

Now, we have to add the 2 half reactions. But not before making sure that the electrons will cancel out. There must absolutely be no net e's in the final balanced equation. By inspection we note that this will only happen if we first multiply the first (oxidation) equation by 5 and the second (reduction) equation by 2, prior to adding both up.

　　　Final balanced REDOX equation:

　　　$5C_2O_4^{2-} + 16H^+ + 2\ MnO_4^- \rightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$　　(Balanced by charge and by mass! Count the charges and the atoms)

5) It is a good habit to start by inspecting the reaction by writing the balanced equation :

　　　$2\ NaOH + H_2SO_4 \rightarrow 2H_2O + Na_2SO_4$;
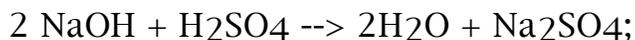
This is a titration problem. If the coefficients were both 1, it would be a simple matter of using $M_1V_1 = M_2V_2$. But it is not so simple, so we can just treat it as a standard stoichiometry problem. At the equivalence point, we have:

[NaOH]=moles of NaOH/L NaOH = moles H_2SO_4 (2 mol NaOH/mol H_2SO_4) / L NaOH

　　　but moles H_2SO_4 = [H_2SO_4] x (L H_2SO_4) = (0.25mol/L)(0.0150L). so, we can write the full solution below:

　.015(.25) mol H_2SO_4 (2mol NaOH /mol H_2SO_4)/0.025L = 0.300 M NaOH.

(sometimes people get confused about what this means. It means that initially, before we even started any titration, the concentration of the NaOH was 0.300M. Titrations are of course done to determine the otherwise unknown concentration of a base or acid solution).

Here is an alternative approach: (this how we reason it out in Chem 201)

At equivalence:　# equiv H_2SO_4 = # equiv NaOH
　　　　　　　2 x #moles H_2SO_4 = #moles NaOH
But moles = MV, so, we can write:

　　　　　　　2 x $M_{H2SO4}V_{H2SO4}$= $M_{NaOH}V_{NaOH}$ => $2M_1V_1 = M_2V_e$ where V_e =
　　vol at equiv

$$M_2 = 2M_1V_1/V_e = 2(15mL)(.250M)/(25 \text{ mL}) = 0.300M \quad \text{(note}$$
that the volume is kept as "mLs" because the units cancel out in the end anyway.)

b) In this question, we have NOT YET REACHED EQUIVALENCE. So, we first calculate how many NaOH's have reacted based on the H_2SO_4 as the limiting reagent. Then, we calculate how much NaOH has remained unreacted by just subtracting it off from the original NaOH. For this problem, it is safest to think in terms of MOLES. Note that this problem is asking what the concentration of NaOH is during the titration. Thus, we need to take into account the actual total volume of the solution after adding H_2SO_4 to it.

Moles NaOH reacted = moles H_2SO_4 added x (2 mol NaOH/1mol H_2SO_4 added)
= (.250 mol H_2SO_4/L)(.0100L)(2)

the abbreviated solution looks like this:

[NaOH]=mol NaOH$_{\text{left over}}$/L total soln = mol NaOH$_i$ -mol NaOH$_{\text{reacted}}$)/(vol NaOH$_i$+vol H$_2$SO$_4$ added) = (M$_1$V$_1$-2M$_2$V$_2$)/(V$_1$+V$_2$)= (.300(25)-2(.25)(10))/(25+10)=.0714 M NaOH

6) (a) Note that the number of moles of H_2SO_4 is determined by using conversion factors. That is then divided by the total final volume (in liters).

[H$_2$SO$_4$]=mol H$_2$SO$_4$/L =25.0 g(1mL sol'n/1.48g sol'n)(18molH$_2$SO$_4$/1000mL)/.500L = 0.608 M

b) OK, don't forget that 5% (by mass) means 5 grams of solute/100 g solution. First you need to determine the concentration of the NaOH solution.

[NaOH] = (5 gNaOH/100g NaOH soln)(1mol NaOH/40g NaOH)(1.1g/mL)(1000mL/L)=1.38 M NaOH;

OK. Let's do the conversion factor method:
Unknown = given x conversion factors

#mL NaOH = (50.0 mL H_2SO_4)(1L/1000mL)(0.608mol H_2SO_4/L H_2SO_4) (2molNaOH/molH_2SO_4)(1L NaOH /1.38 mol NaOH)(1000mL/L) =44.1 mL

Alternative method: At equivalence: # eq H_2SO_4 = # eq NaOH
$$2M_1V_1=M_2V_2 => V_2=2M_1V_1/M_2$$
$$V_2=2(.608M(50.0mL))/1.38M=$$
44.1mLs!

7) a) To get limiting reactant, you can do 2 approaches.

*The "brute force" approach is just to solve for the amount of product you expect based on each one of the reactant amounts given. The reactant that gives the least amount of product is the limiting reactant.

*The other method is to assume one of the reactants is the limiting reactant and then calculate how much of the "other" reactant you will need for this limiting reactant to be completely used up. If your calculation shows that you in fact have more reactant at hand than you need, then your first assumption is correct. Otherwise, it is not and you will need to choose the other reactant as your limiting reactant,

Let's try the second approach: Assume that H_2 is limiting:

Solve for mol N_2 needed to react with the 12.0mol H_2:

#mol N_2 needed = 12.0 mol H_2 x (1 mol N_2/3mol H_2)= 4.0 mol N_2. Since that is less than the 6.0 mol available, N_2 must be in excess by 2.0 mols. H_2 is limiting reactant. #g N_2 excess = 2.0 mol(28.0g/mol)=56.0 g N_2.

b) Theoretical yield of NH_3 = (12.0 mol H_2)(2 mol NH_3/3 mol H_2) =8.0 mol NH_3 (17.0g/mol)=136 g NH_3

c) Recall that % yield = (actual yield/theoretical yield)x100%

=> actual yield = (%yield/100%)(theoretical yield)

So, # mol NH_3= (0.80)(8.0 mol)=6.4 mol.

d) #g NH_3=6.4 mol.x(17.0g/mol)=110 g NH_3

e) # molecules NH_3 = 6.4 mol NH_3 x (6.02×10^{23} molec/mol)= 3.85×10^{24} molecules

f)V_{NH3} = 110g(1L/0.76g)= 140 L

8) Write balanced equation for this redox reaction: (recall how we did it in lecture)

   Cu(s) + 4HNO_3 --> Cu(NO_3)$_2$ + 2NO_2 (g)+ 2H_2O

This is a conversion factor problem:

   #L NO_2 = (0.15cm³Cu)(8.95g Cu /cm³ Cu)(1mol Cu /63.5g Cu)(2molNO_2/molCu) (1L NO_2/2.05g NO_2)(0.70) = 0.014L

9) recall MV=moles; so MV of $H_2C_2O_4$ gives moles $H_2C_2O_4$. From there convert to moles KOH then concentration. But you need balanced eqn: [KOH]=?

    $H_2C_2O_4$ + 2KOH --> $K_2C_2O_4$ + 2H_2O:
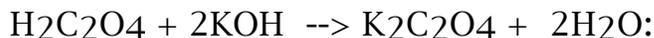
   [$H_2C_2O_4$]=.0200L KOH (.400molKOH/L KOH)(1mol $H_2C_2O_4$ /mol KOH)/.020L $H_2C_2O_4$ =   0.25 M $H_2C_2O_4$

   alternatively, M_1V_1 = 2 M_2V_2 (which is "1"? "1" is KOH in this equation. If you can't clearly determine which is which, don't use this alternative approach. You will have only a 50% chance of succeeding).

      M_2 =M_1V_1/(2V_2) = (.400M)(20.0mL)/(2(16.0mL) = 0.25 M $H_2C_2O_4$