

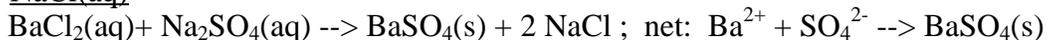
Chem 101 Key for Test #2 Review Guide.

Midterm #2 is on Wednesday, Nov. 19, 2003. Test #2 will focus on Chapters 4-5

Sample short problems:

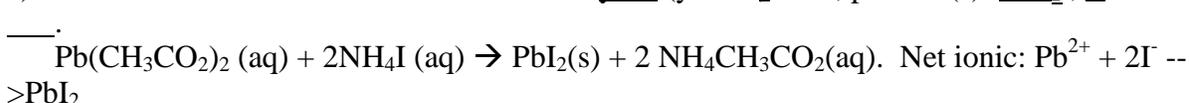
1) answer: #g Na₂CO₃ needed = .250L(.14molNa/L)(1molNa₂CO₃/2molNa)(106g/mol)=1.86g

2) a) sodium sulfate & barium chloride. React? yes (yes/no) product(s): BaSO₄ ppt & NaCl(aq)



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.

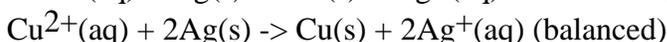
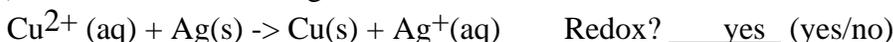
c) Lead acetate & ammonium iodide. React? yes (yes/no) If so, product(s): PbI₂, _



d) sulfuric acid + sodium hydrogen carbonate ? yes (yes/no) If so, product(s): Na₂SO₄, CO₂(g), H₂O(l).



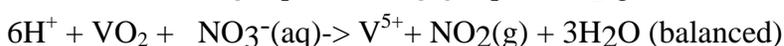
3)) Which of the following are redox reactions? Acid base reactions? **Balance** them anyway.a)



b) metathesis reactions in general. Redox? NO (yes/no)

(Note: metathesis = double replacement = exchange reactions)

c) VO₂(s) + HNO₃(aq) → V(NO₃)₅(aq) + NO₂(g) (unbalanced) Redox? yes (yes/no)



d) CH₄ + O₂ → CO + H₂O (unbalanced) Redox? yes (yes/no)



e) H₂CO₃ + NaCl → Na₂CO₃ + HCl Redox? NO (yes/no) (their oxidation numbers don't change) It's acid base since there is a transfer of protons from H₂CO₃ to Cl⁻.

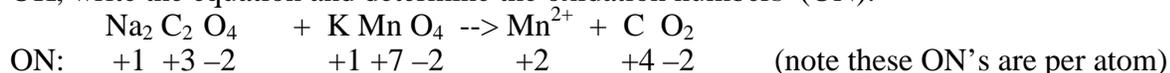
f) H₂CO₃ → H₂O + CO₂? Redox? NO (yes/no) (their oxidation numbers don't change, and no transfer of protons)

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):



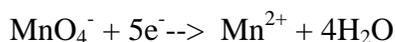
Since the ON of C and of Mn change, we can safely conclude that this is a redox reaction. To balance this, first write the 2 half reactions ("1/2 rxn"). Train yourself to do this in the net ionic form:

Oxidation 1/2 rxn: $\text{C}_2\text{O}_4^{2-} \rightarrow 2\text{CO}_2 + 2\text{e}^-$ (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).

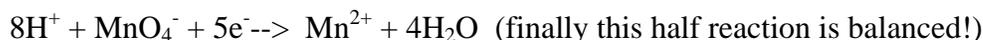
We note that this half reaction equation is ALREADY BALANCED.

Reduction 1/2 rxn: $\text{MnO}_4^- + 5\text{e}^- \rightarrow \text{Mn}^{2+}$ (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:



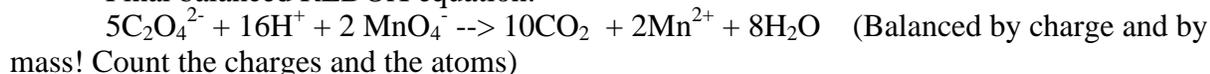
OK, now although the O's are balanced, a new problem arises, we have added 8 H's on the right side. To balance those, we add 8 H^+ 's on the left side:



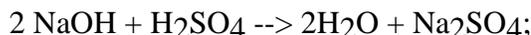
finally, check to see if the charges are balanced. Yes they are! Both sides have a net +2 charge.

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:



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



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:

$$[\text{NaOH}] = \text{moles of NaOH/L NaOH} = \text{moles H}_2\text{SO}_4 (2 \text{ mol NaOH/mol H}_2\text{SO}_4) / \text{L NaOH}$$

but moles $\text{H}_2\text{SO}_4 = [\text{H}_2\text{SO}_4] \times (\text{L H}_2\text{SO}_4) = (0.25\text{mol/L})(0.0150\text{L})$. so, we can write the full solution below:

$$.015(.25) \text{ mol H}_2\text{SO}_4 (2\text{mol NaOH /mol H}_2\text{SO}_4)/0.025\text{L} = 0.300 \text{ M NaOH.}$$

(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
 $2 \times \text{#moles } \text{H}_2\text{SO}_4 = \text{#moles NaOH}$

But moles = MV, so, we can write:

$2 \times M_{\text{H}_2\text{SO}_4} V_{\text{H}_2\text{SO}_4} = M_{\text{NaOH}} V_{\text{NaOH}} \Rightarrow 2M_1 V_1 = M_2 V_e$ where V_e = vol at equiv

$M_2 = 2M_1 V_1 / V_e = 2(15\text{mL})(.250\text{M}) / (25 \text{ mL}) = 0.300\text{M}$ (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 \times (2 mol NaOH/1mol H_2SO_4 added)
 $= (.250 \text{ mol } \text{H}_2\text{SO}_4/\text{L})(.0100\text{L})(2)$

the abbreviated solution looks like this:

$[\text{NaOH}] = \text{mol NaOH}_{\text{left over}} / \text{L total soln} = (\text{mol NaOH}_i - \text{mol NaOH}_{\text{reacted}}) / (\text{vol NaOH}_i + \text{vol } \text{H}_2\text{SO}_4 \text{ added}) = (M_1 V_1 - 2M_2 V_2) / (V_1 + V_2) = (.300(25) - 2(.25)(10)) / (25 + 10) = .0714 \text{ M NaOH}$

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).

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

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.

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

OK. Let's do the conversion factor method:

Unknown = given \times conversion factors

$$\# \text{mL NaOH} = (50.0 \text{ mL H}_2\text{SO}_4)(1\text{L}/1000\text{mL})(0.608\text{mol H}_2\text{SO}_4/\text{L H}_2\text{SO}_4) \\ (2\text{molNaOH}/\text{molH}_2\text{SO}_4)(1\text{L NaOH} /1.38 \text{ mol NaOH})(1000\text{mL}/\text{L}) = 44.1 \text{ mL}$$

Alternative method: At equivalence: # eq H₂SO₄ = # eq NaOH

$$2M_1V_1 = M_2V_2 \Rightarrow V_2 = 2M_1V_1/M_2$$

$$V_2 = 2(.608\text{M}(50.0\text{mL}))/1.38\text{M} = 44.1\text{mLs!}$$

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 H₂ is limiting:

Solve for mol N₂ needed to react with the 12.0mol H₂:

#mol N₂ needed = 12.0 mol H₂ x (1 mol N₂/3mol H₂) = 4.0 mol N₂. Since that is less than the 6.0 mol available, N₂ must be in excess by 2.0 mols. H₂ is limiting reactant. #g N₂ excess = 2.0 mol(28.0g/mol) = 56.0 g N₂.

b) Theoretical yield of NH₃ = (12.0 mol H₂)(2 mol NH₃/3 mol H₂) = 8.0 mol NH₃
(17.0g/mol) = 136 g NH₃

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

$$\Rightarrow \text{actual yield} = (\% \text{yield}/100\%)(\text{theoretical yield})$$

So, # mol NH₃ = (0.80)(8.0 mol) = 6.4 mol.

d) #g NH₃ = 6.4 mol.x(17.0g/mol) = 110 g NH₃

e) # molecules NH₃ = 6.4 mol NH₃ x (6.02x10²³ molec/mol) = 3.85x10²⁴ molecules

f) V_{NH₃} = 110g(1L/0.76g) = 140 L

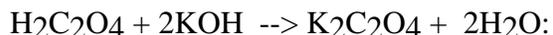
8) Write balanced equation for this redox reaction: (recall how we did it in lecture)



This is a conversion factor problem:

$$\#L \text{ NO}_2 = (0.15\text{cm}^3\text{Cu})(8.95\text{g Cu /cm}^3 \text{ Cu})(1\text{mol Cu /63.5g Cu})(2\text{molNO}_2/\text{molCu}) (1\text{L NO}_2/2.05\text{g NO}_2)(0.70) = 0.014\text{L}$$

9) recall $MV = \text{moles}$; so MV of $\text{H}_2\text{C}_2\text{O}_4$ gives moles $\text{H}_2\text{C}_2\text{O}_4$. From there convert to moles KOH then concentration. But you need balanced eqn: $[\text{KOH}] = ?$



$$[\text{H}_2\text{C}_2\text{O}_4] = .0200\text{L KOH} (.400\text{molKOH/L KOH})(1\text{mol H}_2\text{C}_2\text{O}_4 / \text{mol KOH}) / .020\text{L H}_2\text{C}_2\text{O}_4 = 0.25 \text{ M H}_2\text{C}_2\text{O}_4$$

alternatively, $M_1V_1 = 2M_2V_2$ (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).

$$M_2 = M_1V_1 / (2V_2) = (.400\text{M})(20.0\text{mL}) / (2(16.0\text{mL})) = 0.25 \text{ M H}_2\text{C}_2\text{O}_4$$

10) heat needed = heat needed to melt the ice + heat needed to heat the water from 0°C to 100°C + heat needed to boil the water at 100°C
 $= m\Delta H_{\text{fus}} + mC(\text{Tf}-\text{Ti}) + m\Delta H_{\text{vap}} = m(\Delta H_{\text{fus}} + C\Delta T + \Delta H_{\text{vap}}) =$
 $(12.0\text{g})(333\text{J/g} + (4.18\text{J/molK})(100.\text{K}) + 2260\text{J/g}) = 36,100 \text{ J}$

11) $2x(a) - 6x(b)$: $\Delta H = 2(-1118.4) - 6(-272) = -605 \text{ J}$