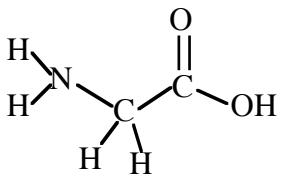
Chemistry 112, Spring 2006 Prof. Metz Final Exam Name _____
Each question is worth 4 points, unless otherwise noted

1. The predominant intermolecular attractive force in solid sodium is:
   (A) metallic
   (B) ionic
   (C) covalent
   (D) dipole-dipole
   (E) induced dipole-induced dipole

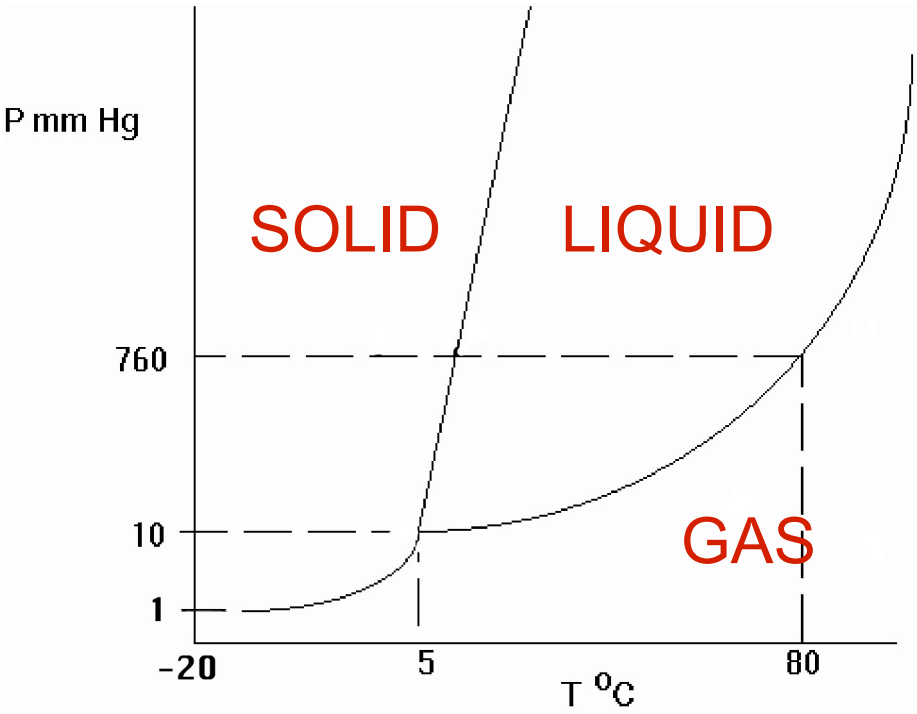
   Sodium is a metal, so it has metallic bonding



2. Glycine (shown above) would be predicted to be most soluble in which solvent:
   (A) CH_4 (B) CCl_4 (C) CCl_3H (D) C_3H_8 *(E) H_2O*

   Glycine is a hydrogen bond donor and acceptor and is polar, so it would be most soluble in a polar, hydrogen-bonding solvent: H_2O

Use the following phase diagram for a substance to answer question 3.



3. A sample of the substance is placed in a container at a temperature of 0°C and a pressure of 500 mm Hg. The pressure is then gradually lowered to 0.5 mm Hg, at constant temperature. What phase changes does the sample go through?

(A) liquid only
(B) solid to gas only
(C) solid to liquid only
(D) liquid to gas only
(E) solid to liquid to gas

Start at 0°C and 500 mm Hg, so it's a solid. Drop the pressure (at constant T), so draw a vertical line down on the phase diagram, end up in the gas region (at 0.5 mm Hg and 0°C), but the substance goes directly from solid to gas (it was never a liquid – it sublimes)

4. 50.00 g of ethylene glycol (MW= 62.0 g/mol), a non-electrolyte, is dissolved in 500.0 g water (MW= 18.00 g/mol) to give a solution with a volume of 550.0 mL. Assume that the solution is ideal. K_{fp} for water is -1.86 °C/m.
The freezing point of this solution in °C is:
(A) –4.59° *(B) –3.00°* (C) –2.73° (D) –1.86° (E) –0.96°

Moles solute = moles ethylene glycol = 50 g glycol/(62.1 g/mol) = 0.805 moles ethylene glycol
Molality of solution = moles solute/kg solvent = 0.805 moles ethylene glycol/0.500 kg water = 1.61 m
$\Delta T_{fp} = K_{fp}\, m_{solute}\, i$
Ethylene glycol is a non-electrolyte, so i=1
ΔT_{fp} = (-1.86 °C/m) (1.61 m) (1) = -3.00 °C
The normal freezing point of water is 0 °C, so the freezing point of this solution is 0 + (-3.00) °C = -3.00 °C

5. Compared to pure water, at 300 Kelvin the osmotic pressure of a 0.010 M solution of NaCl is :
(Note: assume that the NaCl solution is ideal)

(A) 0.49 atm
(B) 0.25 atm
(C) 0.050 atm
(D) 0.025 atm
(E) 0.013 atm

c = concentration = 0.010 M, and i = 2 (NaCl solution has Na+ and Cl- ions, so 2 moles of ions for every more of NaCl)
$\Pi = c\, R\, T\, i$
Π = (0.010 M) (0.0821 L atm/(mol K)) (300 K) (2) = 0.49 atm

6. The reaction
$2\, I^-(aq) + Br_2(aq) \rightarrow I_2(aq) + 2\, Br^-(aq)$ was studied at 25°C. The following results were obtained where

$$\text{Rate} = -\frac{d[Br_2]}{dt}$$

$[I^-]_0$ (mol/L)	$[Br_2]_0$ (mol/L)	Initial rate (mol/L s)
0.080	0.040	12.60×10^{-3}
0.040	0.040	6.30×10^{-3}
0.080	0.020	3.15×10^{-3}

The rate equation for this reaction is:

(A) Rate = $k[I^-]$

(B) Rate = $k[I^-][Br_2]$

(C) Rate = $k[I^-]^2[Br_2]$

(D) Rate = $k[I^-][Br_2]^2$

(E) Rate = $k[I^-]^2[Br_2]^2$

Expt 2 vs 1: $[Br_2]$ unchanged; $[I^-]$ doubles -> rate doubles, so first order in $[I^-]$
Expt 3 vs 1: $[I^-]$ unchanged; $[Br_2]$ doubles -> rate quadruples, so second order in $[Br_2]$
$R = k[I^-][Br_2]^2$

7. The half life for the radioactive decay of ^{14}C is 5720 years.
If a 1.000 g sample of ^{14}C were to sit for 8000 years, how much ^{14}C would remain?

(A) 0.969 g
(B) 0.715 g
(C) 0.489 g
(D) 0.379 g
(E) 0.247 g

$t_{1/2} = 0.693/k$, so
$k = 0.693/(t_{1/2}) = 0.693/(5720 \text{ years}) = 1.2115 \times 10^{-4} \text{ year}^{-1}$
[R] = [R]0 e-kt
$[R] = (1.000 \text{ g})\, e^{-(1.2115 \times 10\text{-}4 \text{ year-}1)(8000 \text{ years})} = 1.00 \text{ g } e^{-0.96923} = 0.379 \text{ g}$

8. A proposed mechanism for the reaction of NO_2 with CO to produce NO and CO_2 is:

$NO_2 + NO_2 \rightarrow NO_3 + NO$ slow
$NO_3 + CO \rightarrow NO_2 + CO_2$ fast

The rate law consistent with this mechanism is:

(A) Rate = $k[NO_2]$
(B) Rate = $k[NO_2]^2$
(C) Rate = $k[NO_2][CO]$
(D) Rate = $k[NO_2]^2[CO]$
(E) Rate = $k[NO_3][NO]/[NO_2]^2$

The slow step is rate-determining, so
Rate = $k[NO_2][NO_2] = k[NO_2]^2$

Use the following thermodynamic information to answer question 9.

Species	**ΔH_f^0(298 K) kJ/mol**	**S^0(298 K) J/(K mol)**	**ΔG_f^0(298 K) kJ/mol**
$CH_4(g)$	-74.87	186.26	-50.8
$O_2(g)$	0	205.07	0
$CO_2(g)$	-393.51	213.74	-394.36
$H_2O(g)$	-241.83	188.84	-228.59

9. Calculate ΔG_{rxn} in kJ/mol at a temperature of 298 K for the reaction

$$CH_4\,(g) + 2\,O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

(A) –572 (B) –623 (C) –674 *(D) –801* (E) –902

$\Delta G_{rxn} = \Sigma\Delta G_f(\text{products}) - \Sigma\Delta G_f(\text{reactants})$
$\Delta G_{rxn} = -394.36 + 2(-228.59) - [-50.8 + 2(0)] = -800.74$ kJ/mol

10. The reaction

$$NO\ (g) + \frac{1}{2} O_2\ (g) \rightarrow NO_2\ (g)$$

has ΔH_{rxn} = -57 kJ/mol and ΔS_{rxn} = -73 J/(mol K).

What is ΔG_{rxn} in kJ/mol at 100 °C for this reaction?

(A) -30 (B) -50 (C) -64 (D) -72 (E) -84

$\Delta G = \Delta H - T\Delta S$
ΔG = -57 kJ/mol – (373 K) (-73 x 10^{-3} kJ/(mol K)) = -29.8 kJ/mol

11. The reaction

$N_{2(g)} + 3\ H_{2(g)} \rightleftharpoons 2\ NH_{3(g)}$ has K_c=11 at 600 K.

A 1 liter flask is filled with 0.01 moles of N_2, 0.03 moles of H_2 and 0.02 moles of NH_3.
Will any reaction occur ? If so, is NH_3 produced or consumed ?

(A) No reaction will occur
(B) A reaction will occur; NH_3 will be consumed
(C) A reaction will occur; NH_3 will be produced

$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Q = $(0.02)^2/[(0.01)(0.03)^3]$ = 1481
Q > K_c, so as the reaction proceeds, Q will become smaller (eventually reaching K_c at equilibrium).
For Q to decrease, need to make less NH_3, and more N_2 and H_2

12. An aqueous solution has a pH of 4.62. The $[OH^-]$ in the solution is

(A) 2.2 x 10^{-1} (C) 1.0 x 10^{-7} *(E) 4.2 x 10^{-10}*
(B) 2.4 x 10^{-5} (D) 8.3 x 10^{-8}

pH = -log$[H_3O^+]$, so
$[H_3O+]$ = 10-pH = $10^{-4.62}$ = 2.40 x 10^{-5}
$[OH-]$ = $10^{-14}/[H_3O+]$ = $10^{-14}/(2.40 \times 10^{-5})$ = 4.16 x 10^{-10}

13. The pH of a 0.150 M solution of formic acid, HCOOH is
(K_a (formic acid) = 1.8 x 10^{-4}).

(A) 0.82 (C) 3.74 (E) 5.38
(B) 2.28 (D) 4.57

$HCOOH + H_2O <-> H_3O^+ + COOH^-$ $K_a = 1.8 \times 10^{-4}$

Initial	*0.150*	*0*	*0*
Change	*-x*	*x*	*x*
Equilibrium	*0.150-x*	*x*	*x*

So, $K_a = [H_3O^+][COOH^-]/[HCOOH] = (x)(x)/(0.150-x) = 1.8 \times 10^{-4}$
Assume x is small, so 0.15-x = 0.15, then
$(x^2)/0.15 = 1.8 \times 10^{-4}$
$x^2 = 2.7 \times 10^{-5}$
$x = [H_3O^+] = 5.2 \times 10^{-3}$ (and note that x is much less than 0.15)
pH = -log $[H_3O^+]$ = 2.28

14. The pH of a 0.05 M solution of nitric acid, HNO_3 is

(A) 0.05 (C) 0.70 (E) 2.00
(B) 0.10 *(D) 1.30*

HNO_3 is a strong acid, so it ionizes completely to form $H^+ + NO_3^-$. The H^+ combines with water to form H_3O^+. So, the H_3O^+ concentration is 0.05 M
pH = $-\log[H_3O^+] = -\log(0.05) = 1.30$

15. HBrO reacts with water via

$$HBrO + H_2O \rightleftharpoons BrO^- + H_3O^+$$

In this acid-base reaction, _____ and _____ act as Bronsted acids; while _____ and _____ are bases:

(A) HBrO, H_2O; BrO^-, H_3O^+

(B) HBrO, BrO^-; H_2O, H_3O^+

(C) HBrO, H_3O^+; H_2O, BrO^-

(D) H_2O, H_3O^+; HBrO, BrO^-

(E) BrO^-, H_3O^+; HBrO, H_2O

HBrO is an acid, it donates a proton to H_2O (base), producing BrO^- (conjugate base) and H_3O^+ (acid)

Question 16 refers to the following gas phase equilibrium for which K_c = 12 at 1100K and $\Delta H° = -198$ kJ/mol.

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$$

16. Addition of $SO_{2\,(g)}$ to an equilibrium mixture of the three gases at constant volume and temperature would cause:

(A) K to increase and the amount of $O_{2\,(g)}$ to decrease.
(B) K to decrease and the amount of $O_{2\,(g)}$ to increase.
(C) K to decrease and the amount of $O_{2\,(g)}$ to decrease.
(D) no change in K but a decrease in the amount of $O_{2\,(g)}$.
(E) no change in K but an increase in the amount of $O_{2(g)}$.

No temperature change, so no change in K.
Add reactant, so reaction proceeds in the forward direction, consuming SO_2 (and O_2) and producing SO_3.

17. 1 mole of $NH_4Cl_{(s)}$ is put into an evacuated 1 liter container at 550 K, and the following reaction occurs:

$$NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$$

At equilibrium, $[NH_{3(g)}] = 2.2 \times 10^{-3}$ M. What is K_c for the reaction ?

(A) 2.4×10^{-6} *(B) 4.8×10^{-6}* (C) 9.6×10^{-6} (D) 1.9×10^{-5} (E) 2.2×10^{-3}

1 mol of HCl is produced for every mole of NH_3 produced, so $[HCl_{(g)}] = [NH_{3(g)}] = 2.2 \times 10^{-3}$ M
$Kc = [NH_3][HCl] = (2.2 \times 10^{-3})(2.2 \times 10^{-3}) = 4.84 \times 10^{-6}$

Questions 18 and 19 refer to a solution made by dissolving 0.010 mol of benzoic acid ($HC_7H_5O_2$) and 0.020 moles of sodium benzoate ($NaC_7H_5O_2$) in enough water to make 1.00 L of solution. pK_a for $HC_7H_5O_2$ is 4.20.

18. What is the approximate pH of this solution?

(A) 4.20 *(C) 4.50* (E) 3.10
(B) 3.90 (D) 2.96

pH=pKa + log([conjugate base]/[acid])
pH=4.20 + log(0.020/0.010)
pH = 4.20 + 0.30 = 4.50

19. What is the approximate pH after the addition of 0.010 mol of HCl to the solution (assume no volume change)?

(A) 2.00 (C) 4.50 (E) 2.96
(B) 3.90 (D) 4.20

Start with 0.010 moles benzoic acid and 0.020 moles benzoate ($C_7H_5O_2^-$) (Initial)
The 0.010 moles of HCl ionize completely to Cl^- and H^+, which combines with water to form H_3O^+, then H_3O^+ reacts with the base: (Change)

$H_3O^+ + C_7H_5O_2^- \rightarrow HC_7H_5O_2 + H_2O$

	$HC_7H_5O_2$	$C_7H_5O_2^-$
Initial	0.010 moles	0.020 moles
Change	+0.010 moles	-0.010 moles
Final	0.020 moles	0.010 moles

pH=pKa + log([conjugate base]/[acid])
pH=4.20 + log(0.010/0.020)
pH = 4.20 - 0.30 = 3.90

20. Calculate the $[ClO^-]/[HClO]$ ratio necessary to give a buffer with a pH = 8.00.
K_a for HClO is 3.5×10^{-8}
(A) 1.00 (B) 1.07 *(C) 3.50* (D) 0.286 (E) 0.932

$Ka = 3.5 \times 10^{-8}$, so pKa = -log(Ka) = 7.46
pH =pKa + log([conjugate base]/[acid])
8.00 = 7.46 + log([conjugate base]/[acid])
log([conjugate base]/[acid]) = 0.54
[conjugate base]/[acid] = $10^{0.54}$ = 3.47

21. What is the $[Ca^{2+}]$ in a saturated $CaCO_3$ solution?
K_{sp} ($CaCO_3$) = 8.7×10^{-9}

(A) 2.1×10^{-3} M (D) 8.7×10^{-9} M
(B) 9.3×10^{-5} M (E) 4.4×10^{-9} M
(C) 6.2×10^{-7} M

$CaCO_{3(s)} \leftrightarrow Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$

Initial	0	0
Change	x	x
Equilibrium	x	x

$K_{sp} = [Ca^{2+}][CO_3^{2-}]^2 = (x)\ (x) = 8.7 \times 10^{-9}$
$x^2 = 8.7 \times 10^{-9}$
$x = [Ca^{2+}] = 9.33 \times 10^{-5}$ M

22. The reaction
$2\ Al_{(s)} + 6\ H^+_{(aq)} \rightarrow 2\ Al^{3+}_{(aq)} + 3\ H_{2(g)}$ has a cell potential E^o = 1.66 V under standard conditions (1 M concentrations for solutions, 1 atmosphere pressure for gases).
What is the cell potential E when $[H^+]$ = 0.01 M, $[Al^{3+}]$ = 0.5 M and H_2 is at a pressure of 1 atmosphere?

(A) 1.64 V (D) 1.23 V
(B) 1.58 V (E) 0.99 V
(C) 1.55 V

$E = E^o - (0.0257/n) \ln(Q)$
$Q = [Al^{3+}]^2 [H_2]^3 / [H^+]^6$
Where $[H_2]$ is the H_2 pressure in atmospheres. Note that $Al_{(s)}$ isn't in Q as it's a solid.
$Q = (0.5)^2 (1)^3 / (0.01)^6 = 2.5 \times 10^{11}$
n = # moles of e- transferred, which is 6
The half reactions are $2\ Al \rightarrow 2\ Al^{3+} + 6e^-$ and $6\ H^+ + 6\ e^- \rightarrow 3\ H_2$
$E = 1.66\ V - (0.0257/6) \ln(2.5 \times 10^{11}) = 1.66 - 0.11 = 1.55\ V$

23. The reaction
$Fe_{(s)} + 2\ H^+_{(aq)} \rightarrow Fe^{2+}_{(aq)} + H_{2(g)}$ has a cell potential $E^o = 0.44$ V under standard conditions. What is ΔG^o for the reaction, in kJ/mol?
(A) -21.2 *(D) -84.9*
(B) -42.4 (E) -96.2
(C) -73.1

$\Delta G^0 = -nFE^0$
n = number of electrons transferred = 2
$F = 9.6485 \times 10^4$
$\Delta G^0 = -nFE^0 = -(2)(9.6485 \times 10^4)(0.44) = -84907$ J/mol = -84.9 kJ/mol

24. Balance the following redox reaction in acidic solution
(Some of the coefficients will be zero !)
(8 pts)
___ $H_3AsO_{4(aq)}$ + ___$Cr_{(s)}$ + ___ H_2O + ___ $H^+_{(aq)} \rightarrow$
___ $H_3AsO_{3(aq)}$ + ___$Cr^{3+}_{(aq)}$ + ___ H_2O + ___ $H^+_{(aq)}$

Oxidation numbers: As (in H_3AsO_4)=+5, Cr (in Cr) = 0; As (in H_3AsO_3) = +3, Cr (in Cr^{3+}) = +3
Write out half reactions, and add e- to balance oxidation numbers
$H_3AsO_4 + 2e^- \rightarrow H_3AsO_3$ (add 2e- to reactants, as As goes from +5 to +3)
$Cr \rightarrow Cr^{3+} + 3e^-$ (add 3e- to products, as Cr goes from 0 to +3)
Add H^+ to balance overall charge, then add H_2O to balance hydrogens and oxygens
$H_3AsO_4 + 2e^- + 2H^+ \rightarrow H_3AsO_3 + H_2O$
$Cr \rightarrow Cr^{3+} + 3e^-$ (no need to add anything, charge and atoms already balanced)
Number of electrons aren't the same, so multiply first reaction by 3 and second reaction by 2 to make them the same. This gives
$3H_3AsO_4 + 6e^- + 6H^+ + 2Cr \rightarrow 3H_3AsO_3 + 3H_2O + 2Cr^{3+} + 6e^-$
Simplify: $3H_3AsO_4 + 6H^+ + 2Cr \rightarrow 3H_3AsO_3 + 3H_2O + 2Cr^{3+}$

Check to make sure that charge and atoms are balanced (reactants vs products)

25. There was also an Extra Credit question worth 5 points

Note: Many questions on this final exam were on your exams 1-3. Your final exam will include some questions from exams 1-3, but not as many.

Additional questions from Chapter 20 that were on Exam 4 in Spring 2006.
Your final exam will include questions on this material.

Questions 26 and 27 refer to the following cell:

A galvanic cell is constructed in which one half cell consists of a Cu electrode in a 1.0 M $CuSO_4$ solution and the other half cell consists of a Ag electrode in a 1.0 M $AgNO_3$ solution.

For $Cu^{2+} + 2e^- \rightarrow Cu$ the standard reduction potential $E^o = +0.34$ V

For $Ag^+ + e^- \rightarrow Ag$ the standard reduction potential $E^o = +0.80$ V

The net cell reaction is:

$$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$$

26. The anode (negative electrode) reaction is:

(A) $Ag^+ + e^- \rightarrow Ag$
(B) $Cu^{2+} + 2e^- \rightarrow Cu$
(C) $Ag \rightarrow Ag^+ + e^-$
(D) $Cu \rightarrow Cu^{2+} + 2e^-$
(E) $Ag^+ \rightarrow Ag^{2+} + e^-$

The two reactions are
$Cu \rightarrow Cu^{2+} + 2e^-$ and $Ag^+ + e^- \rightarrow Ag$
The copper reaction produces electrons, so it occurs at the negative electrode (anode)

27. The cell voltage is:

(A) 1.94 V
(B) 1.14 V
(C) 0.80 V
(D) 0.46 V
(E) 1.24 V

The reactions are

$Ag^+ + e^- \rightarrow Ag$ $E = E^o = +0.80$ V (don't multiply by 2!)

$Cu \rightarrow Cu^{2+} + 2e^-$ $E = -E^o = -0.34$ V (oxidation, so change sign of reduction potential)

Total = sum = 0.46 V

28. In the reaction

$$MnO_4^- + Cu_2O \rightarrow MnO_3^- + CuO$$

(A) Mn is oxidized. Its oxidation number changes from +5 to +3
(B) Mn is reduced. Its oxidation number changes from +5 to +3
(C) Mn is oxidized. Its oxidation number changes from +7 to +5
(D) Mn is reduced. Its oxidation number changes from +7 to +5
(E) Mn is not oxidized or reduced. Its oxidation number is unchanged

In MnO_4^-, Mn has an oxidation number of +7, in MnO_3^- it is +5.
The oxidation number is reduced.

When balancing reactions, please simplify coefficients as much as possible
If a compound doesn't appear as a reactant or a product, use a coefficient of zero.
Write all solutions on the answer sheet !

29. Balance the following redox reaction
(4 pts)

$$2\ Ag^+_{(aq)} + 1\ Cu_{(s)} \rightarrow 2\ Ag_{(s)} + 1\ Cu^{2+}_{(aq)}$$

Assign oxidation numbers, balance half reactions, then combine:
Oxidation numbers: $Ag^+ = +1$, $Cu = 0$; $Ag = 0$, $Cu^{2+} = +2$

$Ag^+ + 1e- \rightarrow Ag$
$Cu \rightarrow Cu^{2+} + 2\ e-$

Multiply first (Ag) reaction by 2 to get e- to balance

30. Balance the following redox reaction in basic solution
(Some of the coefficients will be zero !)
(4 pts)

$NiO_2 + Zn_{(s)} + 2\ H_2O + 0\ OH^-_{(aq)} \rightarrow Ni(OH)_2 + Zn(OH)_2 + 0\ H_2O + 0\ OH^-_{(aq)}$

Oxidation numbers: Ni (in NiO_2) = +4, Zn = 0; Ni (in $Ni(OH)_2$) = +2, Zn (in $Zn(OH)_2$) = +2

$NiO_2 + 2e- + 2H^+ \rightarrow Ni(OH)_2$ Add H^+ to balance charge

$Zn + 2H_2O \rightarrow Zn(OH)_2 + 2e- + 2H^+$

Add H^+ to balance charge, then add H_2O to balance oxygens

Numbers of electrons transferred in the two reactions are the same, so add the reactions together and simplify coefficients ($2H^+$ in reactants cancel $2H^+$ in products)

$NiO_2 + Zn + 2H_2O \rightarrow Ni(OH)_2 + Zn(OH)_2$

There are no H^+ remaining, so don't need to add OH^- to neutralize them.

31. Balance the following redox reaction in acidic solution
(Some of the coefficients will be zero !)
(8 pts)

$2\ CrO_4^{2-}{}_{(aq)} + 3\ SO_3^{2-}{}_{(aq)} + 1\ H_2O + 4\ H^+_{(aq)} \rightarrow$
$2\ Cr(OH)_3 + 3\ SO_4^{2-}{}_{(aq)} + 0\ H_2O + 0\ H^+_{(aq)}$

Oxidation numbers: Cr (in CrO_4^{2-})=+6, S (in SO_3^{2-}) = +4; Cr (in $Cr(OH)_3$) = +3, S (in SO_4^{2-}) = +6

Write out half reactions, add e- to balance oxidation numbers

$CrO_4^{2-} + 3e- \rightarrow Cr(OH)_3$

$SO_3^{2-} \rightarrow SO_4^{2-} + 2e-$

Add H^+ to balance charge, then add H_2O to balance hydrogens and oxygens

$CrO_4^{2-} + 3e- + 5H^+ \rightarrow Cr(OH)_3 + H_2O$

$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2e- + 2H^+$

Number of electrons aren't the same, so multiply first reaction by 2 and second reaction by 3 to make them the same. This gives

$2CrO_4^{2-} + 10H^+ + 3SO_3^{2-} + 3H_2O \rightarrow 2Cr(OH)_3 + 2H_2O + 3SO_4^{2-} + 6H^+$

Simplify: $2CrO_4^{2-} + 4H^+ + 3SO_3^{2-} + H_2O \rightarrow 2Cr(OH)_3 + 3SO_4^{2-}$