Chemistry 112-2014, Exam #4, **VERSION 2** NAME: _____

1. Which of the following will create a buffer solution when equal volumes of each solution are mixed? Choose all that apply—there may be more than one correct answer.

(a) 0.20 M HNO_3 + 0.10 M H_2CO_3
(b) 0.20 M HCO_2H + 0.10 M $NaHCO_2$
(c) 0.20 M HF + 0.20 M CaF_2
(d) 0.10 M HNO_3 + 0.30 M $NaCH_3CO_2$
(e) 0.20 M $NaNO_2$ + 0.10 M NH_3

2. Which of the following acid-base pairs would be used to create a buffer of pH = 9.8, and which species would be present in higher concentration?

Weak Acid	Conjugate Base	K_a	pK_a
$HC_2O_4^-$	$C_2O_4^{2-}$	6.4×10^{-5}	4.19
$H_2PO_4^-$	HPO_4^{2-}	6.2×10^{-8}	7.21
HCO_3^-	CO_3^{2-}	4.8×10^{-11}	10.32

a) $HC_2O_4^-/C_2O_4^{2-}$ $HC_2O_4^-$ in greater concentration
b) $HC_2O_4^-/C_2O_4^{2-}$ $C_2O_4^{2-}$ in greater concentration
c) $H_2PO_4^-/HPO_4^{2-}$ $H_2PO_4^-$ in greater concentration
d) $H_2PO_4^-/HPO_4^{2-}$ HPO_4^{2-} in greater concentration
e) HCO_3^-/CO_3^{2-} HCO_3^- in greater concentration
f) HCO_3^-/CO_3^{2-} CO_3^{2-} in greater concentration

3. Calculate the pH after 0.0200 mol of $NaC_6H_5CO_2$ are added to 175 mL of a 0.180 M solution of $C_6H_5CO_2H$.

a) 3.82 b) 4.00 c) 4.20 d) 4.40 e) 5.86

4. Under which circumstances will $PbCl_2$ be most soluble?

a) in pure water
b) in a solution of 0.10 M $Pb(NO_3)_2$
c) in a solution of 0.10 M NaCl
d) none of these- $PbCl_2$ is insoluble

5. What is the solubility of AgBr, *in grams per liter*? Molar mass = 187.8 g/mol.

a) 1.38×10^{-4} g/L b) 1.01×10^{-10} g/L c) 5.07×10^{-11} g/L

d) 7.35×10^{-7} g/L e) 1.38×10^{-4} g/L

6. What is the molar solubility of CaF_2 in a 0.50 M solution of KF?

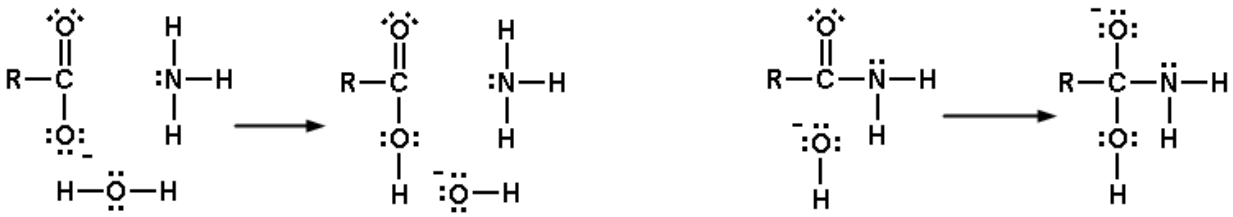
a) 5.3×10^{-11} M b) 1.06×10^{-10} M c) 2.12×10^{-10} M d) 2.37×10^{-4} M e) 0.25 M

7. Which of the following can water (H_2O) not act as?

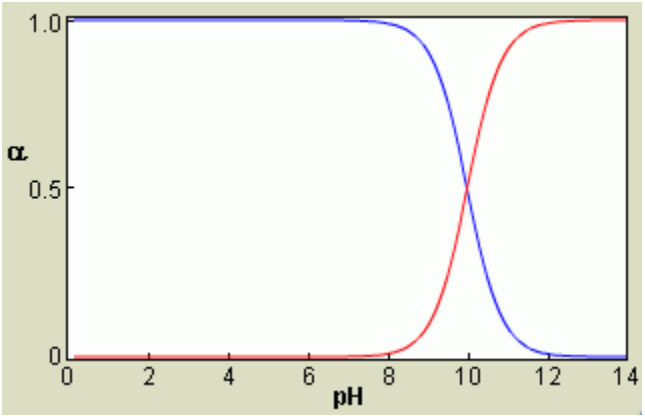
a) Bronsted acid b) Bronsted base c) Lewis acid d) Lewis base

8. Consider the following pair of reaction steps. Add the following labels to the species they represent. Point to the species and add the label. You label four things in total. If something is both Lewis and Bronsted, label it Bronsted. For example, find something that is acting as a Lewis Acid and draw and arrow to it labeled with "LA."

BA = Bronsted Acid LA = Lewis Acid
BB = Bronsted Base LB = Lewis Base



9. Consider the alpha plot below, which is for an acid-base system HA/A^-.



a. which species is mainly prevalent at pH = 8?

HA or A^- or about the same amount of each

b. What is pK_a of HA? _____

c. What is K_a of HA? _____

d. If this acid-base pair were a pH indicator, would it be suitable for the titration of NH_4^+ with NaOH?

Yes or No

10. A saturated solution of chromium(III) hydroxide, $Cr(OH)_3$, is found to have a pH of 10.63. What is K_{sp} for $Cr(OH)_3$ dissolution?

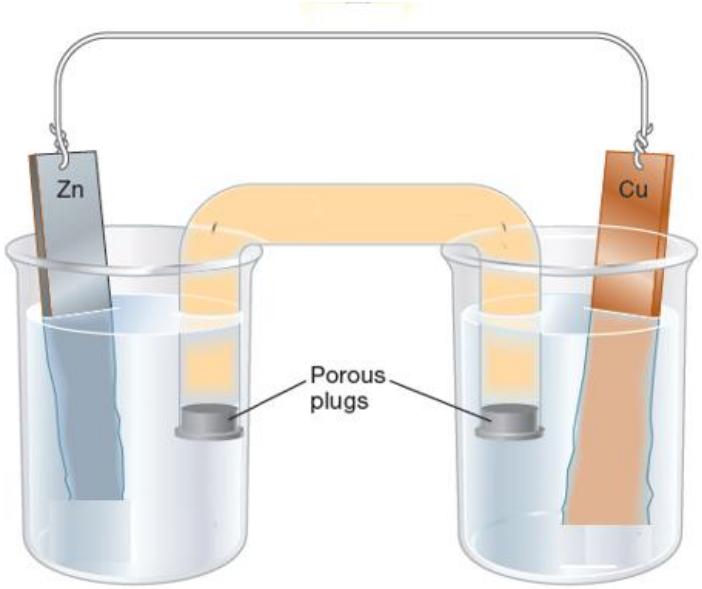
11. The reaction below takes place in an electrochemical cell using compartments containing $Zn/ZnSO_4(aq)$ and $Cu/CuSO_4(aq)$.

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

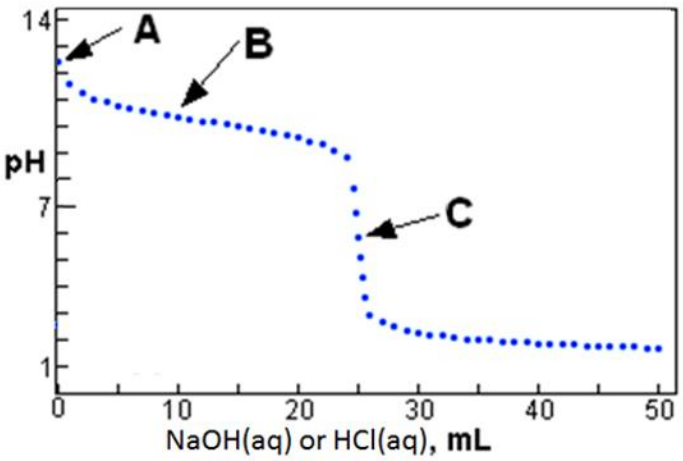
a) write the oxidation half-reaction:

b) write the reduction half-reaction:

c) label the following: anode, cathode, salt bridge, direction of electrons, direction of sulfate ions



12.
Consider the titration curve below involving a weak acid (HA) and a weak base (A^-).



a. Which does this titration represent? Circle it.

weak acid being titrated with NaOH or

weak base being titrated with HCl

b. What is the approximate value of K_a for the acid form (HA) of the studied acid-base pair?

K_a = _____

c. What acid-base species is/are mainly in solution at point A? _____

d. What acid-base species is/are mainly in solution at point B? _____

e. What acid-base species is/are mainly in solution at point C? _____

13. a) What is the pH of a buffer composed of 1.0 L of 0.30 M HCO_2H and 0.50 M HCO^{2-}?

b) What is the pH of this buffer after 0.080 mol NaOH have been added?

14. The Henderson-Hasselbalch equation has a built-in assumption that is not completely valid. Determine the percent error use of the Henderson-Hasselbalch equation leads to when predicting the pH of a buffer containing 0.010 M HF + 0.010 M NaF.

15. What is the concentration of free Ni^{2+} in a 0.200 M solution of $Ni(NH_3)_6^{2+}$? $K_f = 5.5 \times 10^8$

What is the concentration of Ni^{2+} if NH_3 is added to the solution so $[NH_3] = 0.100$ M?

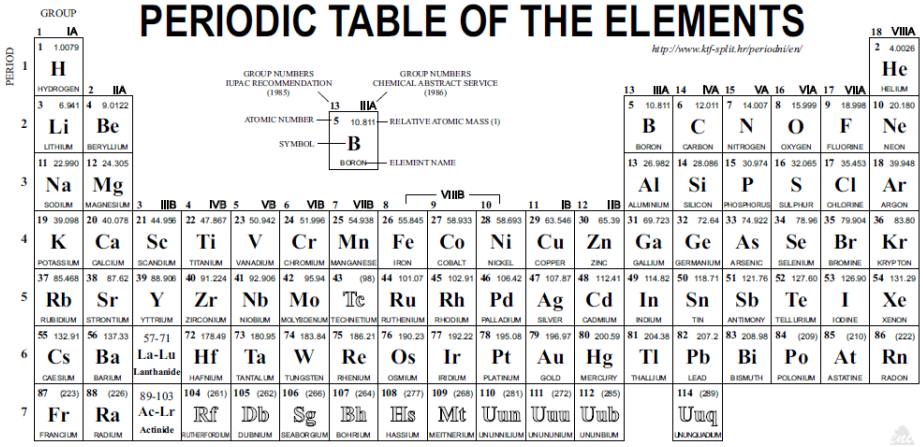
K_a and K_b values for Common Weak Acids and Bases

K_a and K_b Values

Name of Acid	Acid	K_a	Name of Base	Base	K_b
Hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}	sulfate ion	SO_4^{2-}	8.3×10^{-13}
Phosphoric acid	H_3PO_4	7.5×10^{-3}	dihydrogen phosphate ion	$H_2PO_4^-$	1.3×10^{-12}
Hexaaquairon(III) ion	$Fe(H_2O)_6^{3+}$	6.3×10^{-3}	pentaaquahydroxoiron(III) ion	$Fe(H_2O)_5OH^{2+}$	1.6×10^{-12}
Hydrofluoric acid	HF	7.4×10^{-4}	fluoride ion	F^-	1.4×10^{-11}
Formic acid	HCO_2H	1.8×10^{-4}	formate ion	HCO_2^-	5.6×10^{-11}
Benzoic acid	$C_6H_5CO_2H$	6.3×10^{-5}	benzoate ion	$C_6H_5CO_2^-$	1.6×10^{-10}
Acetic acid	CH_3CO_2H	1.8×10^{-5}	acetate ion	$CH_3CO_2^-$	5.6×10^{-10}
Hexaaquaaluminum ion	$Al(H_2O)_6^{3+}$	7.9×10^{-6}	pentaaquahydroxoaluminum ion	$Al(H_2O)_5OH^{2+}$	1.3×10^{-9}
Carbonic acid	H_2CO_3	4.2×10^{-7}	hydrogen carbonate ion	HCO_3^-	2.4×10^{-8}
Hydrogen sulfide	H_2S	1×10^{-7}	hydrogen sulfide ion	HS^-	1×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8}	hydrogen phosphate ion	HPO_4^{2-}	1.6×10^{-7}
Hypochlorous acid	HClO	3.5×10^{-8}	hypochlorite ion	ClO^-	2.9×10^{-7}
Ammonium ion	NH_4^+	5.6×10^{-10}	ammonia	NH_3	1.8×10^{-5}
Hydrocyanic acid	HCN	4.0×10^{-10}	cyanide ion	CN^-	2.5×10^{-5}
Hexaaquairon(II) ion	$Fe(H_2O)_6^{2+}$	3.2×10^{-10}	pentaaquahydroxoiron(II) ion	$Fe(H_2O)_5OH^+$	3.1×10^{-5}
Hydrogen carbonate ion	HCO_3^-	4.8×10^{-11}	carbonate ion	CO_3^{2-}	2.1×10^{-4}
Hydrogen phosphate ion	HPO_4^{2-}	3.6×10^{-13}	phosphate ion	PO_4^{3-}	2.8×10^{-2}

K_{sp} Values for Some Insoluble Salts

Compound	K_{sp} at 25 °C
$CaCO_3$	3.4×10^{-9}
$SrCO_3$	5.6×10^{-10}
$BaCO_3$	2.6×10^{-9}
$BaSO_4$	1.1×10^{-10}
CaF_2	5.3×10^{-11}
$FeCO_3$	3.1×10^{-11}
$Fe(OH)_2$	4.9×10^{-17}
AgCl	1.8×10^{-10}
AgBr	5.4×10^{-13}
AgI	8.5×10^{-17}
Ag_2CrO_4	1.1×10^{-12}
$PbCl_2$	1.7×10^{-5}
$PbCrO_4$	2.8×10^{-13}
$PbBr_2$	6.6×10^{-6}
$PbSO_4$	2.5×10^{-8}



$$pH = pK_a + \log \frac{[base]}{[acid]}$$