12 (1.) Anions from weak acids will make a solid more soluble

1) $AgCl_{(s)}$ → HCl (strong)

2) $KF_{(s)}$ → HF (weak) ✓

3) $NaNO_{3(aq)}$ → HNO_3 (strong); also $NaNO_{3(aq)}$ is not even a solid

4) $H_2CO_{3(aq)}$ → already an acid so putting this in acid will actually ↓ solubility

5) LiBr (s) → HBr (strong)

- Here's why:

① $KF_{(s)} \rightleftharpoons K^+_{(aq)} + F^-_{(aq)}$ if H^+ was added (acidic soln) then $F^-_{(aq)}$ ions would react to form HF thus dec. the amt of $F^-_{(aq)}$ in solution.

② $F^-_{(aq)} + H^+_{(aq)} \rightleftharpoons HF_{(aq)}$

According to Le Chatelier's Principle, if $F^-_{(aq)}$ ions are decreasing, then eq ① will shift to the right to make more F^- thus increasing solubility of $KF_{(s)}$.

12 (2.) ① $H_3AsO_4 \rightleftharpoons H_2AsO_4^- + H^+$

+ ② $HCN \rightleftharpoons CN^- + H^+$ (reverse)

$H^+ + H_3AsO_4 + CN^- \rightleftharpoons H_2AsO_4^- + HCN + H^+$ cancel out H^+'s

↓

$H_3AsO_4 + CN^- \rightleftharpoons H_2AsO_4^- + HCN$ here is the rxn given in the question. We added the acid dissociation of H_3AsO_4 ① and HCN ② but we needed to flip ② to get the original net ionic equation.

To get K_{eq}, it is simply $K_{eq} = K_{a①} \times \frac{1}{K_{a②}}$ because when you add rxns together, you multiply their K's. $K_{a②}$ is flipped because we needed to reverse ② so you take 1 over the Ka.

- Change the pKa's into Ka's

$K_{a①}: 10^{-2.30} = 5.0 \times 10^{-3}$ so $K_{eq} = 5.0 \times 10^{-3} \times \frac{1}{6.2 \times 10^{-10}} = \boxed{8.1 \times 10^6}$

$K_{a②}: 10^{-9.21} = 6.2 \times 10^{-10}$

since $K_{eq} > 1$, the products will be favored

13 ③ $[Cr(CN)_6]^{3-}$ (+3, −6) so Cr^{3+} has −6 extra electrons

$Cr_{(s)}$: s: [1] d: [1][1][1][1][1] normal e⁻ config of Cr(s)

$Cr^{3+}_{(aq)}$: s: [ ] d: [1][1][1][ ][ ] e⁻ config w/ 3 electrons removed

↓ + 6 electrons

but the 6 CN^- ligands will add in 6 new electrons ←

$[Cr(CN)_6]^{3-}$: [1↓] [1↓][1↓][1][1][1]

these are 3 unpaired e⁻

12 ④ 0.10M CH_3COOH

0.05M $Ba(CH_3COO)_2$ → also have 0.10M (2x) of CH_3COO^- ions since there are 2 CH_3COO^- in $Ba(CH_3COO)_2$

Since CH_3COOH and its conjugate CH_3COO^- are both 0.10M, this is @ ½ equivalence so pH = pKa Ka of CH_3COOH is 1.76×10^{-5}

so $pK_a = -\log(K_a) = 4.75$

pH = 4.75

14 ⑤ Since the ratios of cations : anions in the solids are not the same, we can't just compare Ksp's. Instead we solve for each cation's molar solubility then compare.

1) $Ca_3PO_4(s) \rightarrow [3S]^3[S] \Rightarrow 108S^5 = 2 \times 10^{-29}$ ($[3S]^3$: Ca^{2+}, $[S]$: PO_4^{3-})

↳ $S = 7.1 \times 10^{-7}$ then ×3 to get molar solubility of Ca^{2+} ②

$[Ca^{2+}] = [3S] = 2.14 \times 10^{-6}$

2) $CdS(s) \rightarrow [S][S] \Rightarrow S^2 = 8 \times 10^{-27}$

④ $[Cd^{2+}] = [S] = 8.94 \times 10^{-14}$

3) $Al(OH)_3(s) \rightarrow [S][3S]^3 \rightarrow 27S^4 = 1.3 \times 10^{-33} \Rightarrow S = 2.6 \times 10^{-9} \Rightarrow [Al^{3+}] = [S]$ ③

4) $Fe(OH)_2(s) \rightarrow [S][2S]^2 \rightarrow 4S^3 = 1.8 \times 10^{-15} \Rightarrow S = 7.66 \times 10^{-6} \Rightarrow [Fe^{2+}] = [S]$ ①

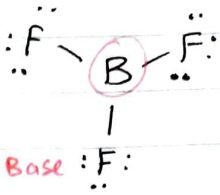
- so now we compare the molar solubilities of the cations from greatest to least:

$Fe(OH)_2(s) > Ca_3PO_4(s) > Al(OH)_3(s) > CdS(s)$

1 (6.) Lewis bases have a lone pair to donate/share with a Lewis Acid.

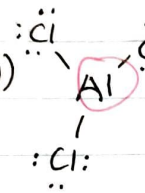
1) 

either LP can be donated to form a new bond → Lewis Base

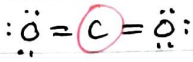
2) 

Boron has space to accept a LP since it is an exception to octet rule → Lewis Acid

3) Ca^{2+} has no LP's to donate but will gladly accept one → Lewis Acid

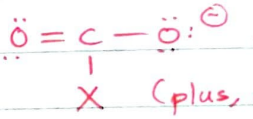
4) 

similar to BF_3 in that Al will accept a LP and form a new bond → Lewis Acid

5) CO_2 

Lewis Acid

interestingly, the C will accept a LP and still maintain octet rule if the = at one of the oxygens becomes a single bond



for example (plus, carbon has no LP's to donate)

4 (7.) I. $\underset{\text{(weak acid) WA}}{CH_3COOH} + OH^- \rightleftharpoons \underset{\text{(conj. base) CB}}{CH_3COO^-} + H_2O(\ell)$

↳ but, this is not a buffer despite the WA and CB being present because the moles of NaOH is greater than moles of CH_3COOH.

$(0.050L)(0.10M) = 0.005$ moles OH^-

$(0.030L)(0.10M) = 0.003$ moles CH_3COOH ⟩ greater than

↳ since the OH^- is greater than CH_3COOH, that means all of our WA will react and disappear and thus cannot be a buffer since buffers need WA + CB

✓ II. $HCl + \underset{WB}{NH_3} \rightleftharpoons \underset{CA}{NH_4^+} + Cl^-$

$(0.050L)(0.10M) = 0.005$ mol HCl

$(0.200L)(0.10M) = 0.020$ mol NH_3 ↑ greater

Since moles of WB > moles HCl, we will still have some WB and CA present. ✓

✓ III. We add $HClO_2$ (WA) and $\underbrace{NaClO_2}_{CB}$, which already has its CB, ClO_2^-, in it so putting these two together will immediately make a buffer.

13 (8.) Both mixtures have HF (WA) and F^- (CB) in them, the difference is the amount that is present.

For buffer ①: $(0.050L)(0.50M) = 0.025$ moles HF

$(0.050)(0.50M) = 0.025$ moles F^-

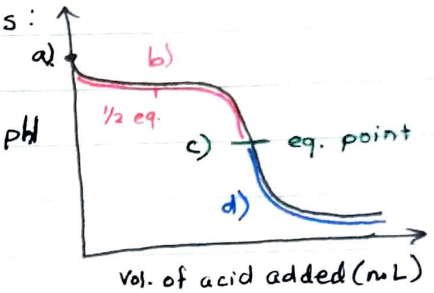
For buffer ②: $(0.200L)(0.50M) = 0.100$ moles HF

$(0.050L)(0.50M) = 0.025$ moles F^-

- what we see is that both buffers can handle the same amount of acid that is thrown at them because the moles of the CB, F^-, are the same.
- Since buffer ② has more moles of WA, HF, than ①, buffer ② has a better buffer capacity if a base was added in.
- The questions asks about if an acid was added so both buffers have the same capacity in that respect.

(9.) see next page...

(9.) We have a weak base (NO_2^-) - strong acid (HBr) titration which means the titration curve will look something like this:



a) Before you add any acid, only NO_2^- is reacting w/ water so to find pH just do an ICE table.

$$NO_2^-(aq) + H_2O(l) \rightleftharpoons HNO_2(aq) + OH^-$$

	NO_2^-		HNO_2	OH^-
I	0.0989 M *		0	0
C	$-x$		$+x$	$+x$
E	$0.0989 - x$ (use 5% rule to ignore $-x$)		x	x

$$\frac{x^2}{0.0989} = 1.4 \times 10^{-11} \quad \text{(5\% rule)}$$

$x = [OH^-] = 1.18 \times 10^{-6}$ M

$pOH = -\log[OH^-] = 5.93$

then use $pH + pOH = 14$ to find $pH = 8.07$

* We weren't given conc of NO_2^- but we can find it since we know it took 88.07mL of HBr to get to the eq. point. At the eq. point the moles of our WB = moles of SA and vice versa. So: ~~$(0.1123\text{M HBr})(0.100\text{L}) = 0.01123\text{mol} = (x\text{ M } NO_2^-)($~~

$$(0.1123\text{ M HBr})(0.08807\text{L}) = 0.00989\text{ mol} = (x\text{ M } NO_2^-)(0.100\text{L})$$

$$x = 0.0989\text{ M } NO_2^-$$

b) Next since we are @ half eq. $pOH = pK_b$, then we can find pH

$pOH = -\log[1.4 \times 10^{-11}] = 10.85 \rightarrow pH = 14 - 10.85 = \boxed{3.15}$

- if you are before the eq. point but NOT at the ½ eq. point, then you do another ICE table, but with new concentrations of NO_2^- and HBr because the vol. has changed, and then use the Henderson-Hasselbach equation to find pOH to get pH

c) When you reach the eq. point, all the NO_2^- is reacted with and its conjugate HNO_2 was made and that reacts with water. So do an ICE table with the conjugate to get pH.

$$HNO_2(aq) + H_2O(l) \rightleftharpoons NO_2^-(aq) + H_3O^+(aq)$$

	HNO_2		NO_2^-	H_3O^+
I	0.0526 M *		0	0
C	$-x$		$+x$	$+x$
E	$0.0526 - x$ (5% rule)		x	x

$$\frac{x^2}{0.0526\text{ M}} = 7.14 \times 10^{-4}$$

$x = [H^+] = 8.13 \times 10^{-3}$ M

$pH = -\log(H^+) = \boxed{2.21}$

* remember our conc. will Δ since our volumes added together. So: $[HNO_2] = \frac{(0.100\text{L})(0.0989\text{M})}{(0.100\text{L} + 0.08807\text{L})} = 0.0526\text{ M}$

d) now we are past the eq. point so we need to find how much acid HBr is in excess.

$(0.100\,L\ HBr)(0.1123\,M\ HBr) = 0.01123\,mols\ H^+$ initially

$(0.100\,L\ NO_2^-)(0.0989\,M\ NO_2^-) = 0.00989\,mols\ NO_2^-$ initially

$$\frac{(0.01123\,mols\ H^+ - 0.00989\,mols\ NO_2^-)}{(0.100\,L + 0.100\,L)} = 6.7\times10^{-3}\,M\ H^+ \text{ in excess}$$

$-\log(6.7\times10^{-3}\,M) \Rightarrow \boxed{2.17 = pH}$

5

10. Rules of naming:

- alphabetical order for ligands
- use di-, tri-, tetra-, etc. for ligands except for "en" which uses bis-, tris-, tetrakis-, etc.
  - ↳ * these prefixes do not affect abc order
- only add "-ate" to metal if the thing in brackets is the anion of the complex.

Solution:

all the answers are correct in terms of naming but platinum should be (IV) not (V)

$$[\overbrace{Pt^{+4}(NH_3)_4^{0}Br^{-1}Cl^{-1}}^{2+}](NO_3)_2^{2-}$$

[4] (11.) To find which precipitates first we must compare molar solubilities of Fe^{2+} in both $FeCO_3$ and $Fe(OH)_2$ because the ratios of cations to anions are not the same.

$[Fe^{2+}][CO_3^{2-}] = 1.0\times10^{-13} \Rightarrow [Fe^{2+}] = 1\times10^{-10}\ M$

(CO_3^{2-}: 0.001M)

$[Fe^{2+}][OH^-]^2 = 1.0\times10^{-15} \Rightarrow [Fe^{2+}] = 1\times10^{-9}\ M$

(OH^-: 0.001M)

now we see that $FeCO_3$ will precip. first and that we only need 1×10^{-10}M $Fe(NO_3)$ to precip.

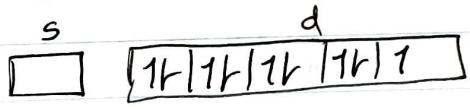
[2] (12.) Nonmetal hydrides (HX and HY)

- more EN that X or Y is, the shorter the bond between H–X and H–Y and thus less acidic since the H cannot come off as easily.
- so ↑ EN => smaller bond => less acid
  ↓ EN => larger bond => more acidic

Oxoacids (HXO_2 and HYO_2)

- if # of oxygens are same, then acidity ↑ with ↑ EN
- if # of oxygens are diff, then the greater # of oxygens is more acidic
  ↳ so: X is more EN than Y; H–X bond is shorter than H–Y

[2] (13.) Cu^{2+}



paramagnetic ✓ (has at least 1 unpaired e^-)

H_2O ligand → weak ✓ → high spin ✓

so I, III, IV

(14.) A complex ion is colorless when the cation metal is diamagnetic

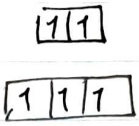
5

I: $[FeCN_6]^{4-}$ Fe^{2+}:

colorless ✓

CN^- is strong ligand (low spin) diamagnetic

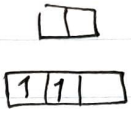
II: $[Mn(H_2O)_6]^{2+}$ Mn^{2+}:



colored ✗

H_2O is weak ligand (high spin)

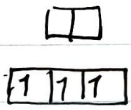
III: $[Ti(NO_2)_6]^{4-}$ Ti^{2+}:



colored ✗

NO_2^- is strong ligand (low spin)

IV: $[V(H_2O)_6]^{2+}$ V^{2+}:



colored ✗

H_2O is weak ligand (high spin)

V: $[Zn(H_2O)_6]^{2+}$ Zn^{2+}:

colorless ✓

H_2O is weak ligand (high spin)

I, V

there's a typo in the review document, answer (5) should say I, V

(1:2 $CaCl_2$ & 1:2 $CuCl_2$)

1 (15.) Since the cation:anion ratios are the same, we can know Cu^{2+} will precip. before Ca^{2+} since the Ksp is smaller. So to "separate" the metal ions we can precip Cu^{2+} but not the Ca^{2+} ions by finding [NaOH] right before Ca^{2+} will begin to precip.

$$[Ca^{2+}][OH^-]^2 = 6.5\times10^{-6} \Rightarrow \boxed{[OH^-] = 4.0\times10^{-3}\,M}$$

(Ca^{2+} = 0.40M)

~~(16.) $CoS(s) \rightleftharpoons Co^{2+}(aq) + S^{2-}(aq)$~~

~~$Co^{2+}(aq) + [Co(OH)_4]$~~

1 (16.)

$$CoS(s) \rightleftharpoons Co^{2+}(aq) + S^{2-}(aq) \qquad K_{sp}$$

$$+\ Co^{2+}(aq) + 4OH^-(aq) \rightleftharpoons [Co(OH)_4]^{2-}(aq) \qquad K_f$$

$$CoS(s) + 4OH^-(aq) \rightleftharpoons [Co(OH)_4]^{2-}(aq) + S^{2-}(aq) \qquad K_{eq} = K_{sp}\cdot K_f =$$

I	2.0M	0	0
C	−4S	+S	+S
E	2−4S (5% rule ignore −4S)	S	S

$$K_{eq} = \frac{[Co(OH)_4]^{2-}_{(aq)}[S^{2-}]}{[OH^-]^4}$$

"S" represents the new molar solubility of CoS(s)

$$2\times10^{-11} = \frac{S^2}{(2.0M)^4} \Rightarrow \boxed{S = 1.8\times10^{-5}\,M}$$

when its placed in NaOH. The molar solubility changes because the Co^{2+} ions are decreased to form $[Co(OH)_4]^{2-}$ and so according to Le Châtelier's, more CoS(s) will have to dissolve.

4 (17.) (1) and (5) are just ionic compounds so they won't be basic.
(2) has NH_4^+ which is an acid so not gonna make a basic solution.
(3) and (4) both have conjugated bases (CH_3COO^- and CN^- respectively) but which one is a stronger base? Use K_b's to determine.

$$K_b(CH_3COO^-) = \frac{1\times10^{-14}}{1.76\times10^{-5}} = 5.7\times10^{-10}$$

$$K_b(CN^-) = \frac{1\times10^{-14}}{6.17\times10^{-10}} = 1.6\times10^{-5} \text{ so } \boxed{CN^- \text{ is stronger base}}$$