CHEM 123 SAMPLE EXAM QUESTIONS

1. How many spheres lie within one unit cell of a body-centred cubic lattice of identical, hard spheres?

A. one
[✓ B. two]
C. four
D. eight
E. nine

SC = 1
BCC = 2
FCC = 4

2. Aluminum (Al) crystallizes in a face-centred cubic structure. If the radius of the aluminum atom is 140 pm, then what is edge length of the unit cell?

A. 200 pm
B. 280 pm
C. 320 pm
[✓ D. 400 pm]
E. 560 pm

SC → $a = 2R$
BCC → $a = \frac{4}{\sqrt{3}}R$
FCC → $a = 2\sqrt{2}R = 2\sqrt{2}(140\text{pm}) = 396\text{pm}$
$= 400\text{pm}$

☆ 3. Which one of the following statements is correct for a closest-packed lattice of identical, hard spheres?
cubic closest packed = FCC

A. The coordination number of each sphere in the lattice is eight. × C.N. = 12
B. The number of tetrahedral holes in the lattice is exactly equal to the number of spheres. × 4 spheres & 8 tetrahedral holes
[✓ C. There are more tetrahedral holes than octahedral holes in the lattice.] ✓ 8 tetr. vs. 4 oct.
D. More than 80% of the total volume of the lattice is occupied by spheres. × NO → 74%
E. The unit cell of the lattice is either simple cubic or face-centred cubic. ✓ / NO.

☆ 4. The overall chemical equation for the combustion of carbon monoxide, CO, is

$$2\,CO(g) + O_2(g) \rightarrow 2\,CO_2(g).$$

Which one of the following expressions is correct?

A. $d[CO]/dt = d[CO_2]/dt$
B. Rate = $k[CO]^2\,[O_2]$
C. Rate = $[CO]^2/([CO]^2\,[O_2])$
[✓ D. $d[CO_2]/dt = -2\,d[O_2]/dt$]
E. $d[CO]/dt + d[O_2]/dt = d[CO_2]/dt$

$$-\frac{1}{2}\frac{d[CO]}{dt} = -\frac{d[O_2]}{dt} = \frac{1}{2}\frac{d[CO_2]}{dt}$$

$$\frac{d[CO_2]}{dt} = -2\frac{d[O_2]}{dt}$$ ✓

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$\rightarrow$ 1st order.'

5. In a certain experiment, the rate law for the reaction, $H_2O_2 \rightarrow$ products, was determined to be $-d[H_2O_2]/dt = k\,[H_2O_2]$ with $k = 1.06\times10^{-3}$ min^{-1}. What is the concentration of H_2O_2 after 100 min, if the initial concentration of H_2O_2 was 0.200 mol L^{-1}?

A. 0.0212 mol L^{-1}
B. 0.0940 mol L^{-1}
C. 0.106 mol L^{-1}
D. 0.157 mol L^{-1}
[✓ E.] 0.180 mol L^{-1}

$$[H_2O_2]_t = [H_2O_2]_0 e^{-kt}$$

$$[H_2O_2]_{100} = (0.200\,M)\, e^{-1.06\times10^{-3}\,min^{-1}(100\,min)}$$

$$[H_2O_2]_{100} = 0.180\,M$$

6. The decomposition of compound A follows a second-order rate law. Which one of the following plots is linear? ([A] is the concentration and *t* is time.)

A. [A] vs. *t*
B. ln[A] vs. *t*
[✓ C.] 1/[A] vs. *t*
D. [A] vs. 1/*t*
E. ln[A] vs. 1/*t*

Zero: [A] vs t is linear with slope $-k$

First: ln[A] vs. t is linear with slope $-k$

second: $\frac{1}{[A]}$ vs. t is linear with slope k

7. The activation energy (E_a) for a certain reaction is 182 kJ/mol and the rate constant (k) is 1.57×10^{-5} mol^{-1} L s^{-1} at 973 K. What is the value of the rate constant at 1073 K?

A. 1.79×10^{-5} mol^{-1} L s^{-1}
[✓ B.] 1.28×10^{-4} mol^{-1} L s^{-1}
C. 1.93×10^{-5} mol^{-1} L s^{-1}
D. 1.28×10^{-2} mol^{-1} L s^{-1}
E. 7.80×10^{-4} mol^{-1} L s^{-1}

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{-E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \rightarrow$$

$$k_2 = k_1\, e^{-E_a/R\,(1/T_2 - 1/T_1)}$$

$$k_2 = (1.57\times10^{-5}\,mol^{-1}L\,s^{-1})\, e^{\frac{-182000\,J\,mol^{-1}}{8.3145\,J\,mol^{-1}K^{-1}}\left(\frac{1}{1073\,K} - \frac{1}{973\,K}\right)} = 1.28\times10^{-4}\,mol^{-1}L\,s^{-1}$$

8. The reaction of NO with O_2 to form NO_2 contributes to the formation of urban smog. The following mechanism has been proposed:

$\rightarrow$ overall rxn: $2NO + O_2 \rightarrow 2NO_2$

$$NO + NO \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} N_2O_2 \quad \text{(fast equilibrium)} \quad k_1, k_{-1} \text{ are } \underline{\text{big}}$$
$$N_2O_2 + O_2 \xrightarrow{k_2} 2\,NO_2 \quad \text{(slow)} \quad k_2\ \underline{\text{small}}$$

What is the rate law derived from this mechanism?

[✓ A.] Rate = k $[NO]^2\,[O_2]$
B. Rate = k $[NO]^2$
C. Rate = k $[NO]\,[O_2]^2$
D. Rate = k $[NO]\,[O_2]$
E. Rate = k $[O_2]^2$

intermediate

① RDS : Rate = $k_2[N_2O_2][O_2]$

② SSA for $[N_2O_2]$: Rate consumed = Rate produced

$$\underbrace{k_2[N_2O_2][O_2]}_{\approx 0} + k_{-1}[N_2O_2] = k_1[NO]^2$$

small compared to

$$\rightarrow [N_2O_2] = \frac{k_1}{k_{-1}}[NO]^2$$

③ Subst. ② into ① $\rightarrow$ Rate $= \frac{k_1k_2}{k_{-1}}[NO]^2[O_2] = k[NO]^2[O_2]$

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9. In a certain reaction (at 298 K), the concentration of compound X, in mol L^{-1}, varies with time, in seconds, as follows:

time:	0	2	4	6	10	20
[X]:	10.0	8.19	6.70	5.49	3.68	1.35

① zero: $k_1 = \frac{[A]_0 - [A]_t}{t} = \frac{10-8.19}{2} = 0.905$

$k_2 = \frac{10-6.70}{4} = 0.825$ × not zero.

Which one of the following statements is correct?

② First: $k_1 = -\frac{1}{t}\ln\left(\frac{[A]}{[A]_0}\right) = -\frac{1}{2}\ln\left(\frac{8.19}{10}\right) = 0.0998$

$k_2 = -\frac{1}{4}\ln\left(\frac{6.70}{10}\right) = 0.100$ ≈ same FIRST

A. The reaction is zero-order with respect to X. ×
[✓ B.] The reaction is first-order with respect to X.
C. The reaction is second-order with respect to X.
D. The rate constant k decreases as the reaction time increases. × constant (except with temp. or ionic strength)
E. Compound X is a product of the reaction. × NO! The [X] decreases with time!

10. For the reaction, A + B → C, the following initial rates were obtained. What is the overall order of the reaction?

i.e. a+b=? → Rate = $k[A]^a[B]^b$

	$[A]_0$ (in mol L^{-1})	$[B]_0$ (in mol L^{-1})	Rate (in mol L^{-1} s^{-1})
①	0.020	0.0040	2.5×10^{-5}
②	0.020	0.0080	3.5×10^{-5}
③	0.070	0.0040	8.8×10^{-5}

② ÷ ① $\frac{3.5\times10^{-5}}{2.5\times10^{-5}} = \frac{\cancel{k}\,\cancel{(0.020)^a}(0.0080)^b}{\cancel{k}\,\cancel{(0.020)^a}(0.0040)^b} \rightarrow 1.4 = 2^b$

$\frac{\ln(1.4)}{\ln 2} = b$

$b \approx \frac{1}{2}$

A. one-half
B. one
[✓ C.] one and one-half
D. two
E. two and one-half

③ ÷ ② $\frac{8.8\times10^{-5}}{3.5\times10^{-5}} = \frac{\cancel{k}(0.070)^a(0.0040)^{\frac{1}{2}}}{\cancel{k}(0.020)^a(0.0080)^{\frac{1}{2}}}$

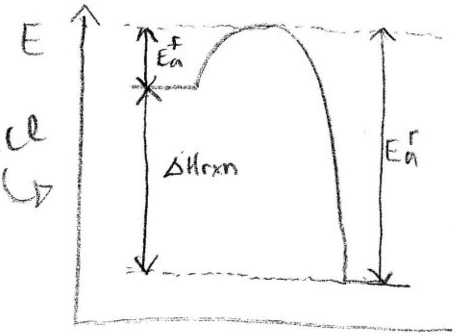
$\rightarrow 2.51 = 3.5^a(0.707) \rightarrow a = \frac{\ln 3.55}{\ln 3.5} \approx 1 \quad \therefore 1 + \frac{1}{2} = 1.5$ ☆

☆ 11. The reaction $NO_2 + CO \rightarrow NO + CO_2$ is an elementary process with $\Delta H^o = -225$ kJ/mol. The activation energy for the forward reaction is 125 kJ/mol. What is the activation energy for the reverse reaction?

exothermic

[✓ A.] 350 kJ/mol
B. 325 kJ/mol
C. 225 kJ/mol
D. 125 kJ/mol
E. 100 kJ/mol

Since ↳



$\therefore E_a^r = E_a^f + \Delta H_{rxn}$

$E_a^r = (125 + 225)$ kJ/mol

$E_a^r = 350$ kJ/mol

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12. An elementary process which contributes to the depletion of atmospheric ozone, O_3, is

$$O_3 + NO \rightleftharpoons NO_2 + O_2 .$$

Let k_{for} and k_{rev} be the rate constants for the forward and reverse processes, respectively. Let K_c be the equilibrium constant. Which one of the following expressions is correct when this process is at equilibrium?

for the equilibrium : Rate Forward ≈ Rate reverse

A. $k_{for} = k_{rev}$
B. $k_{for}\, k_{rev} = K_c$
C. $k_{rev} = 1/k_{for}$
D. $k_{for} + k_{rev} = K_c$
[✓ E.] $k_{for}/k_{rev} = K_c$

$k_{for}[O_3][NO] = k_{rev}[NO_2][O_2]$, rearrange

$$\rightarrow \frac{k_{for}}{k_{rev}} = \frac{[NO_2][O_2]}{[O_3][NO]} = K_c!$$

13. The equilibrium constant for the gas-phase reaction below is K_c=5.10 at 700 K.

$$CO + \underline{H_2O} \rightarrow CO_2 + H_2$$

do not ignore gaseous H_2O!

One mole each of CO, H_2O, CO_2, and H_2 is placed in a 10.0-L flask and the temperature is raised to 700 K. What is the concentration of H_2 when the reaction reaches equilibrium?

A. 0.0400 mol L^{-1}
B. 0.0672 mol L^{-1}
C. 0.103 mol L^{-1}
[✓ D.] 0.139 mol L^{-1}
E. 0.167 mol L^{-1}

	CO	$+ H_2O \rightleftharpoons$	CO_2	$+ H_2$
① i	0.1M	0.1M	0.1M	0.1M
③ c	$-x$	$-x$	$+x$	$+x$
④ e	$0.1-x$	$0.1-x$	$0.1+x$	$0.1+x$

② → check Q!

$$Q_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{0.1^2}{0.1^2} = 1$$

$Q_c < K_c$ (reaction goes right)

⑤ $K_c = 5.10 = \frac{(0.1+x)^2}{(0.1-x)^2} \rightarrow \sqrt{5.10} = \frac{0.1+x}{0.1-x} \rightarrow x = 0.0386M$

$\hookrightarrow [H_2] = (0.1 + 0.0386)M = \underline{0.139\ M}$

14. Consider the equilibrium constant (K_c) for the reaction below.

$$N_2 + 2\,O_2 \rightarrow 2\,NO_2$$

Which one of the following statements is true?

A. K_c changes as the initial concentrations of the reactants change. × NO (only temperature dependendant)
B. K_c changes as the pressure changes. × Same as A.
C. K_c is lowered by the addition of a catalyst. × NO (rxn will come to equilibrium faster only)
D. K_c is raised by the addition of a catalyst. × NO Same as C.
[✓ E.] None of the statements above is true.

15. Select the ion having the lowest concentration in a 0.10 mol L^{-1} aqueous solution of Na_2S. (H_2S is a weak acid with $K_{a1} = 1.0\times10^{-7}$ and $K_{a2} = 1.0\times10^{-19}$.)

$$Na_2S \xrightarrow{100\%} 2Na^+ + S^{2-}$$

	Na_2S	$2Na^+$	S^{2-}
① F	0	0.2M	0.1M

A. $Na^+ \simeq 0.2M$
[✓ B.] $S^{2-} = x$ (very small)
C. $OH^- \approx 0.1M$
D. $HS^- \approx 0.1M$

$K_{b1} = K_w/K_{a2} = 10^{-14}/10^{-19} = 10^5$. (effectively 100% completion)

	S^{2-}	$+ H_2O \rightleftharpoons$	HS^-	$+ OH^-$
i	0.1M	/	0	≈0
c	−0.1M	/	+0.1M	+0.1M
i′	0	/	0.1M	0.1M
c′	x		$-x$	$-x$
e′	x		≈0.1M	≈0.1M

very small!

very small

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16. Which one of the following is the strongest acid?

[✓ A.] Hydrogen iodide, HI
B. Hydrogen bromide, HBr
C. Hydrogen chloride, HCl
D. Hydrogen fluoride, HF
E. Hydrocyanic acid, HCN

} Acidity increases as size of halogen increases down column of the periodic table: $\underbrace{HI > HBr > HCl}_{\text{strong acids.}}$

} weak acids on data sheet!!

strongest acid HI > HBr > HCl > HF > HCN

17. What is $[H_3O^+]$ in a 0.12 mol L^{-1} ascorbic acid, $\overset{H_2A}{H_2C_6H_6O_6}$, solution? Ascorbic acid is a diprotic acid with $K_{a1} = 7.9\times10^{-5}$ and $K_{a2} = 1.6\times10^{-12}$.

[✓ A.] 3.08×10^{-3} mol L^{-1}
B. 4.30×10^{-7} mol L^{-1}
C. 6.70×10^{-7} mol L^{-1}
D. 2.33×10^{-8} mol L^{-1}
E. 3.25×10^{-12} mol L^{-1}

① For a <u>weak</u> polyprotic acid we treat finding the pH or $[H_3O^+]$ the same as a <u>weak</u> monoprotic acid.

	H_2A +	H_2O $\xrightleftharpoons{K_{a1}}$	HA^- +	H_3O^+
i	0.12M	/	0	≈0
c	$-x \leftarrow$ small	/	+x	+x
e	≈0.12	/	x	x

$$K_a = \frac{x^2}{0.12-x} \approx \frac{x^2}{0.12} \rightarrow x = \sqrt{7.9\times10^{-5}(0.12)} = [H_3O^+] = 3.08\times10^{-3}\,M$$

18. How many moles of formic acid, HCOOH, must be dissolved in 1.00 L of pure water to give a solution with a pH of 2.50? Formic acid is a monoprotic acid with $K_a = 1.9\times10^{-4}$.

A. 0.00010 mol
B. 0.0053 mol
[✓ C.] 0.053 mol
D. 0.035 mol
E. 0.0032 mol

	HA +	H_2O $\xrightleftharpoons{K_a}$	A^- +	H_3O^+
i	$n/1.0L$	/	0	≈0
c	$-10^{-2.5}$	/	$+10^{-2.5}$	$+10^{-2.5}$
e	$\frac{n}{1.0L} - 10^{-2.5}$	/	$10^{-2.5}$	$10^{-2.5}$

$$K_a = 1.9\times10^{-4} = \frac{10^{-5}}{n - 10^{-2.5}} \rightarrow n = \frac{10^{-5} + 1.9\times10^{-4}(10^{-2.5})}{1.9\times10^{-4}} = \underline{0.056\text{ mol}}$$

19. Which of the following relationships is true for a 1.0 mol L^{-1} $H_2SO_4(aq)$ solution at 298 K? For H_2SO_4, $K_{a1} \gg 1$ and $K_{a2} = 1.2\times10^{-2}$.

A. $[HSO_4^-] > [H_3O^+] > [SO_4^{2-}]$
[✓ B.] $[H_3O^+] > [HSO_4^-] > [SO_4^{2-}]$
C. $[HSO_4^-] > [SO_4^{2-}] > [H_3O^+]$
D. $[H_3O^+] > [SO_4^{2-}] > [HSO_4^-]$
E. $[SO_4^{2-}] > [H_3O^+] > [HSO_4^-]$

①

	H_2SO_4 +	H_2O $\xrightarrow{K_{a1} = \text{strong}}$	HSO_4^- +	H_3O^+
i	1.0M	/	0	≈0
c	−1.0M	/	+1.0M	+1.0M
F	0	/	1.0M	1.0M

← ②

	HSO_4^- +	H_2O $\xrightleftharpoons{K_{a2}}$	SO_4^{2-} +	H_3O^+
i	1.0M	/	0	1.0M
c	−x	/	+x	+x
e	1.0M − x	/	x	1.0M + x

③ ∴,

$[HSO_4^-] = 1.0M - x$

$[SO_4^{2-}] = x$

$[H_3O^+] = 1.0M + x$

and $[H_3O^+] > [HSO_4^-] > [SO_4^{2-}]$

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20. An aqueous solution contains the following ions. Which one makes the solution basic?

[✓ A.] OBr^- → conjugate base of weak acid ($HOBr$), so weak base ✓
B. Na^+ → not acidic/basic
C. ClO_4^- → conjugate base of strong acid ($HClO_4$), so not basic
D. NO_3^- → conjugate base of strong acid (HNO_3), so not basic
E. K^+ → not acidic/basic

21. The percent ionization of a weak monoprotic acid (HA) in a 0.10 mol L^{-1} solution is 8.60%. What is the percent ionization of HA in a 0.20 mol L^{-1} solution?

A. exactly the same (8.60%)
B. exactly half (4.30%)
C. exactly double (17.2%)
[✓ D.] less than 8.60%
E. greater than 8.60%

☆ Increasing the concentration of an acid decreases its percent ionization, but not by half.

$$f\ NaA \xrightarrow{100\%} Na^+ + A^- \quad (n_{A^-} = n_{NaA}) ☆$$

22. How many moles of sodium acetate, CH_3COONa, must be added to 1.0 L of 0.100 mol L^{-1} acetic acid, CH_3COOH, to give the solution a pH of 4.75?

$K_a = 1.8 \times 10^{-5}$

A. 1.0 mol
B. 0.50 mol
[✓ C.] 0.10 mol
D. 0.010 mol
E. 0.0010 mol

☆ aside pH = pKa here. This is only when $n_{HA_0} = n_{A^-_0}$! ← no calculation required

	HA	+ H_2O	⇌ A^-	+ H_3O^+
i	0.1M	/	$n_{A^-}/1.0L$	≈ 0
c	$- 10^{-4.75}M$	/	$+ 10^{-4.75}M$	$+ 10^{-4.75}M$
e	$\boxed{0.1M - 10^{-4.75}M}$ small!	/	$\boxed{n_{A^-}/1.0L + 10^{-4.75}M}$ small!	$\boxed{10^{-4.75}M}$

$$K_a \approx \frac{[A^-]_0[H_3O^+]_{eq}}{[HA]_0} = \frac{n_{A^-}(10^{-4.75})}{0.1} \rightarrow n_{A^-} = n_{NaA} = \underline{0.1\ mol}$$

23. Separate, saturated aqueous solutions of the hydroxides shown below are prepared at 298 K. Which substance gives a solution with the highest $[OH^-]$? ice table gives

A. $Fe(OH)_3$, $K_{sp} = 1.6\times10^{-39}$
B. $Cr(OH)_3$, $K_{sp} = 6.3\times10^{-31}$
C. $Fe(OH)_2$, $K_{sp} = 7.9\times10^{-16}$
[✓ D.] $Sr(OH)_2$, $K_{sp} = 3.2\times10^{-4}$
E. $Ca(OH)_2$, $K_{sp} = 6.5\times10^{-6}$

A, B: 3:1 salts $x = \sqrt[3]{\frac{K_{sp}}{4}}$

C, D, E: 2:1 salts ice table gives $x = \sqrt{K_{sp}}$

A. $x = [OH^-] = 7.4 \times 10^{-14} M$
B. $x = [OH^-] = 5.4 \times 10^{-11} M$
C. $x = [OH^-] = 2.8 \times 10^{-8} M$
(D.) $x = [OH^-] = 1.8 \times 10^{-2} M$
E. $x = [OH^-] = 2.6 \times 10^{-3} M$

here x increases as K_{sp} increases so $Sr(OH)_2$ yields the highest $[OH^-]$

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24. Equal volumes of two solutions, one containing both 0.010 mol L^{-1} KBr and 0.010 mol L^{-1} KCl and the other containing 0.50 mol L^{-1} $AgNO_3$, are mixed. The solid is removed. What is $[Cl^-]/[Br^-]$ ratio in the solution? For AgCl, $K_{sp} = 1.8\times10^{-10}$ and for AgBr, $K_{sp} = 5.0\times10^{-13}$.

A. 0.0028
B. 1.0
C. 36.0
[✓ D. 360]
E. 3600

① consider each ppte rxns: $Ag^+_{(aq)} + Br^-_{(aq)} \longrightarrow AgBr_{(s)}$

I 0.25M   0.005M   0
C −0.005M   −0.005M   +0.005M
F 0.245M   0   0.005M

then, again for AgCl(s)

so $[Ag^+] = 0.24M$

② for $AgBr_{(s)} \rightleftharpoons Ag^+ + Br^-$

i   0.24   0
c   +x   +x
e $\approx\boxed{0.24}$ $\boxed{x}$

$x = [Br^-] = \frac{ksp}{[Ag^+]} = \frac{5.0\times10^{-13}}{0.24} = 2.08\times10^{-12}M$

and for $[Cl^-] = \frac{ksp}{0.24} = \frac{1.8\times10^{-10}}{0.24} = 7.50\times10^{-10}M \therefore \frac{[Cl^-]}{[Br^-]} = 360$

25. Which one of the following substances would increase the molar solubility of AgCl in aqueous solution?

shift rxn to right

A. NaCl ← adds Cl^- NO!
[✓ B. NH_3] ← reacts with Ag^+ YES!
C. $AgNO_3$ ← adds Ag^+ NO!
D. HCl ← adds Cl^- NO!
E. CH_3COOH ← NO!

$AgCl_{(s)} \rightleftharpoons Ag^+ + Cl^-$ ← anything that removes Ag^+ or Cl^- shifts the reaction to the right (dissolving $AgCl_{(s)}$)

26. What is the concentration of CH_3COO^- ($\rightarrow A^-$) at equilibrium when 0.10 mol of CH_3COOH is dissolved in 1.00 L of 0.10 mol L^{-1} HCl solution at 298 K? For CH_3COOH, $K_a = 1.76\times10^{-5}$.

COMMON ION EFFECT

A. 0.010 mol L^{-1}
B. 1.76×10^{-3} mol L^{-1}
C. 1.76×10^{-4} mol L^{-1}
[✓ D. 1.76×10^{-5} mol L^{-1}]
E. 1.76×10^{-6} mol L^{-1}

① $HCl + H_2O \xrightarrow{100\%} H_3O^+ + Cl^-$

F 0   /   0.1M   0.1M

② $HA + H_2O \overset{Ka}{\rightleftharpoons} A^- + H_3O^+$

i 0.1M   /   0   0.1M
c −x   /   +x   +x ← small
e ≈0.1M   /   x   ≈0.1M

③ $Ka \approx \frac{x(0.1)}{0.1}$

$x = Ka$
$= 1.76\times10^{-5}M$

27. If $[S^{2-}]$ is maintained at 1.0×10^{-13} mol L^{-1}, what is the maximum possible $[Ag^+]$ in an aqueous solution? For Ag_2S, $K_{sp} = 1.6\times10^{-49}$.

A. 1.6×10^{-36} mol L^{-1}
B. 1.6×10^{-23} mol L^{-1}
C. 6.1×10^{-19} mol L^{-1}
[✓ D. 1.3×10^{-18} mol L^{-1}]
E. 2.6×10^{-18} mol L^{-1}

$Ag_2S_{(s)} \overset{ksp}{\rightleftharpoons} 2Ag^+_{(aq)} + S^{2-}_{(aq)}$

i   /   0 1.0×10^{-13}
c   /   +2x   +x ← very small
e   / $\boxed{2x}$ $\boxed{\approx1.0\times10^{-13}}$

$$K_{sp} \approx (2x)^2(1.0\times10^{-13})$$

$$x = \sqrt{\frac{1.6\times10^{-49}}{4\times10^{-13}}} = 6.32\times10^{-19}M$$

$$[Ag^+] = 2x = 1.3\times10^{-18}M$$

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28. Which one of the following half-reactions would produce the largest cell potential when coupled to the half-reaction $Ag^+ + e^- \rightarrow Ag(s)$ under standard conditions? For $Ag^+(aq) + e^- \rightarrow Ag(s)$, E^o is 0.799 V.

A. $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$, $E^o = -0.74$ V
B. $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$, $E^o = -0.41$ V
C. $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$, $E^o = -0.04$ V
D. $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$, $E^o = -0.236$ V
[✓ E.] $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$, $E^o = -0.762$ V

$E^o_{cell} = E^o_{ox} + E^o_{red}$

Since $E^o_{ox} = -E^o_{red}$, then rxn E. gives the greatest E^o_{cell} →

$E^o_{cell} = -(-0.762V) + 0.799V$
$= \underline{1.56 \text{ V}}$

29. Which one of the following statements is correct?

A. An electrolytic cell makes use of a spontaneous oxidation-reduction reaction. (non-spontaneous) ✗
B. Oxidation occurs at the cathode of an electrochemical cell. ✗ oxidation is anode
C. Reduction is a loss of electrons. ✗ LEO says GER or OIL RIG (gain!)
[✓ D.] Electrons flow from the anode of an electrochemical cell to the cathode. ✓ ALWAYS
E. An electron transfer reaction <u>must be carried</u> out in an electrochemical cell. ✗
↳ doesn't have to be a cell

30. What is the value of *n* when the Nernst equation, $E = E^0 - (0.0592/n) \log Q$, is applied to the reaction, $2 H_2 + O_2 \rightarrow 2 H_2O$?

A. $n = 0$
B. $n = 1$
C. $n = 2$
D. $n = 3$
[✓ E.] $n = 4$

From data sheet:

$(2H^+ + 2\bar{e} \rightleftharpoons H_2) \times 2$ and flip
$O_2 + 4H^+ + \boxed{4\bar{e}} \rightleftharpoons 2H_2O$

$2H_2 + O_2 \rightleftharpoons 2H_2O$

$n = 4\bar{e}$'s

CHEM 123 Exam Data Sheet

Please detach, need not be returned

Physical Constants:

Avogadro's Number $N_A = 6.0221 \times 10^{23}$
Charge of an electron $e = 1.6022 \times 10^{-19}$ C
Faraday constant $F = 96485$ C
Gas constant $R = 0.08205$ L atm mol^{-1} K^{-1} = 8.314 J mol^{-1} K^{-1}
Planck's constant $h = 6.6261 \times 10^{-34}$ J.s
At 298 K, $2.303RT/F = 0.0592$, and $K_w = 10^{-14}$

K_{sp} for Some Salts and K_f for some complexes:

AgCl	1.8×10^{-10}	$Ag_3(PO_4)$	1.8×10^{-18}	MnS	3.0×10^{-13}
AgBr	5.0×10^{-13}	$Al(PO_4)$	9.8×10^{-21}	ZnS	2.5×10^{-21}
AgI	8.3×10^{-17}	$Ca_3(PO_4)_2$	1.3×10^{-32}	CdS	1.0×10^{-27}
$PbCl_2$	1.7×10^{-5}	$Ni_3(PO_4)_2$	4.7×10^{-32}	Ag_2S	1.6×10^{-49}
AgOH	1.5×10^{-5}	Ag_2CrO_4	1.2×10^{-12}	Ba CO_3	2.6×10^{-9}
$Sr(OH)_2$	3.2×10^{-4}	$SrCrO_4$	3.6×10^{-5}	Li_2CO_3	1.7×10^{-3}
$Ca(OH)_2$	6.5×10^{-6}	$BaCrO_4$	1.2×10^{-10}	$MnCO_3$	2.2×10^{-11}
$Mg(OH)_2$	7.1×10^{-12}			Complexes	K_f
$Mn(OH)_2$	6.0×10^{-14}	$Al(OH)_3$	1.3×10^{-33}	$Ag(NH_3)_2^+$	1.6×10^7
$Fe(OH)_2$	7.9×10^{-16}	$Cr(OH)_3$	6.3×10^{-31}	$Cd(NH_3)_4^{2+}$	1.0×10^7
$Zn(OH)_2$	4.5×10^{-17}	$Fe(OH)_3$	1.6×10^{-39}	$Cu(NH_3)_4^{2+}$	1.1×10^{13}

Selected Portion of the Periodic Table of Elements

H																	
1.0																	
Li	Be											B	C	N	O	F	Ne
6.9	9.0											10.8	12.0	14.0	16.0	19.0	20.1
Na	Mg											Al	Si	P	S	Cl	Ar
23.0	24.3											27.0	28.0	31.0	32.1	35.5	40.0
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
39.0	40.0	45.0	47.9	50.9	52.0	54.9	55.8	58.9	58.7	63.5	65.4	69.7	72.6	74.9	79.0	79.0	83.8
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
85.5	87.6	88.9	91.2	92.9	95.9	98	101	103	106	108	112	115	119	122	128	127	131
Cs	Ba	La-	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
133	137	139	178	181	184	186	190	192	195	197	201	204	207	209	209	210	222

Key formulas: At 298 K, $E = E^o - (0.0592/n)\log Q$

$$\ln\frac{k_2}{k_1} = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right); \qquad [A] = [A]_o\, e^{-kt}; \qquad k\, t_{1/2} = \ln 2 = 0.693$$

$$[A]^{-1} - [A]_o^{-1} = kt; \qquad k\, t_{1/2}[A]_o = 1; \qquad K_a = \frac{K_w}{K_b}$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}; \quad \ln\frac{K_2}{K_1} = -\frac{\Delta H^o}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$PV = nRT \qquad K_p = K_c\,(RT)^{\Delta n} \qquad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Some Standard Reduction Potentials:

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	2.889 V
$Au^+(aq) + e^- \rightarrow Au(s)$	1.691
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.512
$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$	1.42
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	1.229
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	1.078
$Ag^+(aq) + e^- \rightarrow Ag(s)$	0.799
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	0.77
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	0.401
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	0.339
$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-(aq)$	0.22
$AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$	0.0732
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.0000
$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$	-0.04
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	-0.127
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$	-0.236
$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$	-0.282
$Tl^+(aq) + e^- \rightarrow Tl(s)$	-0.336
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$	-0.402
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	-0.41
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$	-0.74
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	-0.762

K_a and K_b for some Acids and Bases:

Acid	**K_a**
Acetic, CH_3COOH	1.76×10^{-5}
Chloroacetic, $CH_2ClCOOH$	1.41×10^{-3}
Ascorbic, $H_2C_6H_6O_6$	7.9×10^{-5} (K_{a1})
$HC_6H_6O_6^-$	1.6×10^{-12} (K_{a2})
Benzoic, $HC_7H_5O_2$	6.46×10^{-5}
Formic, HCOOH	1.9×10^{-4}
Hypobromous, HOBr	2.06×10^{-9}
Hypochlorous, HOCl	2.8×10^{-8}
Hydrofluoric, HF	3.53×10^{-4}
Hydrocyanic, HCN	4.93×10^{-10}
Hydrogen sulfide, H_2S	1.02×10^{-7} (K_{a1})
HS^-	1.0×10^{-13} (K_{a2})
Nitrous, HNO_2	6.0×10^{-4}
Phosphoric, H_3PO_4	7.52×10^{-3} (K_{a1})
$H_2PO_4^-$	6.23×10^{-8} (K_{a2})
HPO_4^{2-}	4.5×10^{-13} (K_{a3})
Sulfuric, H_2SO_4	strong
HSO_4^-	1.20×10^{-2} (K_{a2})
Base	**K_b**
Ammonia, NH_3	1.74×10^{-5}
Aniline, $C_6H_5NH_2$	4.3×10^{-10}
Trimethylamine $N(CH_3)_3$	6.3×10^{-5}

☆ WARNING, this data sheet has changed over the years!