ACS Exams

American Chemical Society
Division of Chemical Education
Examinations Institute

General Chemistry, Second-Term Practice Items

Study Question – Chapter 9 (Solutions and Aqueous Reactions, Part 2)

1A. What is the molar concentration of a 2.52 *m* aqueous sugar (sucrose, $C_{12}H_{22}O_{11}$) solution?

molar mass, sucrose	342.34 g·mol^{-1}
density, $C_{12}H_{22}O_{11}(aq)$	1.23 g·mL^{-1}

(A) 1.66 M **(B)** 2.05 M **(C)** 3.10 M **(D)** 3.60 M

$M = \frac{mol}{L\ solution}$

Work:

$\frac{2.52\ mol\ Sucrose}{kg\ solvent} \times \frac{1\ kg\ H_2O}{} = 2.52\ moles$

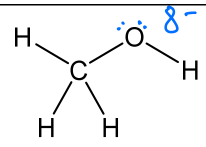
Mass Sucrose: $\frac{2.52\ moles}{} \Big| \frac{342.34\ g}{1\ mol} = 862.7\ g$ Sucrose

total mass $= 1000\ g + 862.7 = \frac{1862.7\ g\ solution}{} \Big| \frac{1\ mL}{1.23\ g} \Big| \frac{1\ L}{1000\ mL} = 1.514\ L$ of solution

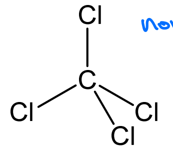
$M = \frac{2.52\ mol}{1.514\ L} = 1.66\ M$

Practice Question – Chapter 9 (Solutions and Aqueous Reactions, Part 2)

1B. Other than dispersion forces, what is the predominant intermolecular force present in a solution of methanol and carbon tetrachloride?



methanol



non-polar

carbon tetrachloride

(A) ion-dipole forces X no ions (charges)

(B) ion-induced dipole forces - no ions...

(C) dipole-dipole forces -only one has a dipole

(D) dipole-induced dipole forces ok... I just call this dispersion...

Work:

References below are to the ACS Exams General Chemistry Study Guide

For more practice, please see: Study Questions: SQ1-4 (p. 114-116)
Practice Questions: PQ1-10 (p. 119-120)

Study Question – Chapter 9 (Solutions and Aqueous Reactions, Part 2)

2A. In order to determine the change in freezing point of 5.0 g of NaCl in 75 g of water, what needs to be included in the calculation?

I. molar mass of NaCl	**II.** molar mass of H_2O
III. freezing point depression constant of NaCl	**IV.** freezing point depression constant of H_2O
V. conversion between g and kg	**VI.** number of ions in NaCl

(A) **I.**, **II.**, and **IV.** **(B)** **I.**, **II.**, **III.** and **IV.** **(C)** **I.**, **IV.**, and **V.** [✓ **(D)** **I.**, **IV.**, **V.** and **VI.**]

Work:

$\Delta T_f = K_f \cdot m \cdot i$ i if ioniz

to get $m = \frac{mol}{kg\ solvent}$ need molar mass of NaCl I.

need K_f ... but not for the solute ✗

V - I guess so...

II - No change in mass ✗

IV - yes K_f

VI - yes $i = ?$ or assume complete dissociati

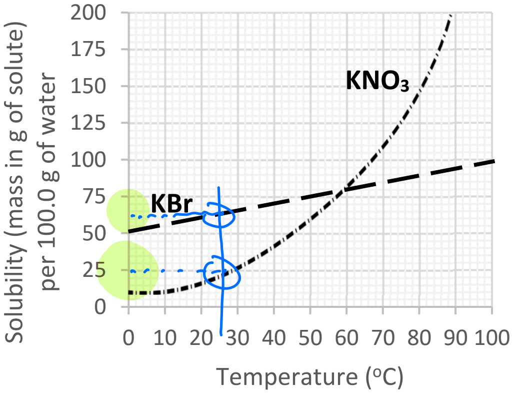
I, IV, V, & VI

Practice Question – Chapter 9 (Solutions and Aqueous Reactions, Part 2)

2B. You have two solutions both at 25°C:

(1) 50.0 g of KBr in 100 g of water
(2) 50.0 g of KNO_3 in 100 g of water

Using the solubility curve shown, which statement is true?



(A) Both are saturated. only KNO_3 at 25°C

(B) Both are unsaturated. No, only KBr

(C) Only KBr(aq) is saturated. ✗

[✓ **(D)**] Only KNO_3(aq) is saturated. ✓

Work:

For more practice, please see: Study Questions: SQ5-8 (p. 117-119)
Practice Questions: PQ11-20 (p. 120-121)

Study Question – Chapter 10 (Kinetics)

3A. A series of experiments were carried out for a reaction:

$$2A + B \rightarrow C$$

Based on the data, what is the rate law?

Exp	$[A]_0$ / M	$[B]_0$ / M	rate / $M \cdot min^{-1}$
1	0.500	0.500	0.0250
2	0.500	1.00	0.0250
3	1.00	0.500	0.0500
4	1.00	1.00	0.0500

(A) rate = k[A] **(B)** rate = $k[A]^2$ **(C)** rate = k[A][B] **(D)** rate = $k[A]^2[B]$

Work: $\frac{rate_2}{rate_1} = \left(\frac{[B]}{[B]}\right)^{☺}$ $= \frac{0.0250}{0.0250} = \left(\frac{1.00}{0.500}\right)^{☺} \rightarrow$ $1 = 2^{☺}$ ☺ = 0 so, $[B]^0$

$\frac{rate_3}{rate_1} = \left(\frac{[A]}{[A]}\right)^{🛸} \rightarrow$ $\frac{0.0500}{0.0250} = \left(\frac{1.00}{0.500}\right)^{🛸} \rightarrow 2 = 2^{🛸}$ 🛸 = 1 so, $[A]^1$

rate = k [A]

Practice Question – Chapter 10 (Kinetics)

3B. There are two reactions with the two rate constants shown to the right. What can you conclude about these reactions?

Reaction 1: $2A \rightarrow$ products	$k = 0.450\ s^{-1}$
Reaction 2: $2B \rightarrow$ products	$k = 0.0875\ s^{-1}$

(A) Reaction 1 is faster than reaction 2. Both reactions are first order.

(B) Reaction 1 is slower than reaction 2. Both reactions are first order.

(C) Reaction 1 is faster than reaction 2. Both reactions are second order.

(D) Reaction 1 is slower than reaction 2. Both reactions are second order.

Work:

rate = k [ ]

$\frac{M}{s} = s^{-1} \cdot M$ so it would be 1st order

bigger k = higher rate for the same order

For more practice, please see: Study Questions: SQ1-2 (p. 124-125)
Practice Questions: PQ1-5 (p. 131)

Study Question – Chapter 10 (Kinetics)

4A. The first-order, radioactive decay of nitrogen-16 is 7.1 seconds. What fraction of a sample of nitrogen remains after 3 half-lifes (or 21.3 seconds) has elapsed?

(A) $\frac{1}{3}$ **(B)** $\frac{1}{6}$ [✓ **(C)** $\frac{1}{8}$] **(D)** $\frac{1}{9}$

Work:

$\frac{1}{2} \rightarrow \frac{1}{4} \rightarrow \frac{1}{8}$

Practice Question – Chapter 10 (Kinetics)

4B. For a reaction, a series of experiments were carried out, and the rate of the reaction was measured over time. The generic reaction is: A → products. Which statement is correct?

(A) If the plot of [A] vs. time is nonlinear, the reaction is zero order.

(B) If plot of [A] vs. time is nonlinear, the reaction is first order.

(C) If plot of ln[A] vs. time is linear, the reaction is zero order.

[✓ **(D)**] If plot of ln[A] vs. time is linear, the reaction is first order.

Work: [A] vs time = linear = zero order

ln[A] vs time = linear for 1st order

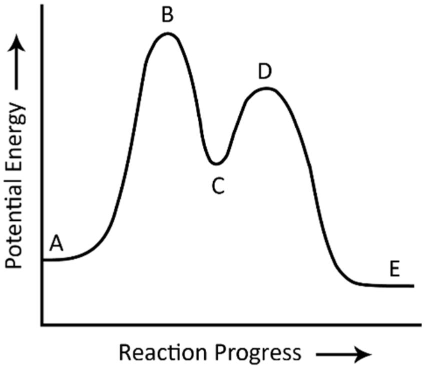
$\frac{1}{[A]}$ vs time = linear for 2nd order

For more practice, please see: Study Questions: SQ3-5 (p. 125-127)
Practice Questions: PQ6-11 (p. 132)

Study Question – Chapter 10 (Kinetics)

5A. Which mechanism is consistent with the energy diagram?

(A)	Step 1: A→E
(B)	Step 1: A→B Step 2: C→D
(C)	Step 1: A→C Step 2: C→E
(D)	Step 1: A→B Step 2: B→C Step 3: C→D Step 4: D→E



Work:

Step1 A → C

Step 2 C → E

C is a reaction intermediate

B & D are activated complexes

overall A → E

Practice Question – Chapter 10 (Kinetics)

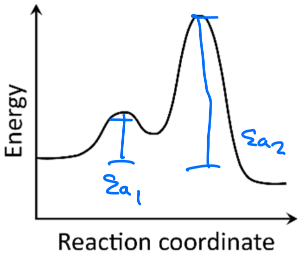
5B. For the energy diagram shown, what are the relationships between the activation energies (E_{a1} and E_{a2}) and the rate constants (k_1 and k_2)?

(A) $E_{a1} < E_{a2}$; $k_1 < k_2$

(B) $E_{a1} < E_{a2}$; $k_1 > k_2$

(C) $E_{a1} > \cancel{E_{a2}}$; $k_1 < k_2$

(D) $E_{a1} \cancel{>} E_{a2}$; $k_1 > k_2$



Work:

$E_{a_1} < E_{a_2}$

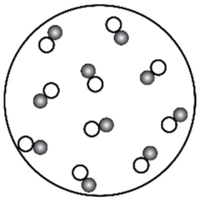
$K_1 > K_2$

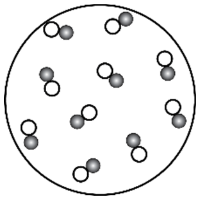
smaller activation energy = faster reaction rates, not as much energy require

For more practice, please see: Study Questions: SQ6-9 (p. 128-130)
Practice Questions: PQ12-30 (p. 132-136)

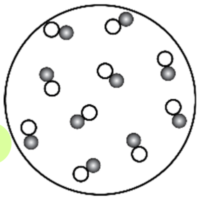
Study Question – Chapter 11 (Equilibrium)

6A. For the reaction of $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$, the equilibrium mixture is show to the right with the equilibrium constant of K_c. What is the value of $K_{c'}$ for the reaction of $2NO(g) \rightleftharpoons N_2(g) + O_2(g)$ and what does this reaction look like at equilibrium?

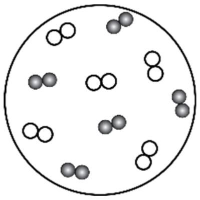


(A) 

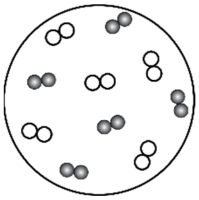
$K_{c'} = -K_c$

(B) 

$K_{c'} = \frac{1}{K_c}$

(C) 

$K_{c'} = -K_c$

(D) 

$K_{c'} = \frac{1}{K_c}$

Work:

Practice Question – Chapter 11 (Equilibrium)

6B. What is the equilibrium expression for the formation of ammonium nitrate, NH_4NO_3?

$$N_2(g) + 2H_2(g) + {}^3/_2O_2(g) \rightleftharpoons NH_4NO_3(s)$$

(A) $K_c = \frac{[NH_4NO_3]}{[N_2][H_2]^2[O_2]^{3/2}}$

(B) $K_c = \frac{1}{[N_2][H_2]^2[O_2]^{3/2}}$

(C) $K_c = \frac{[N_2][H_2]^2[O_2]^{3/2}}{[NH_4NO_3]}$

(D) $K_c = [N_2][H_2]^2[O_2]^{3/2}$

Work:

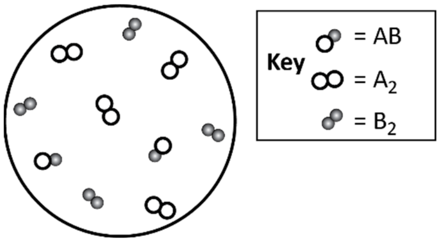
$K_c = \frac{\text{“products”}}{\text{“reactants”}}$ Solids excluded

$K_c = \frac{1}{[N_2][H_2]^2\ [O_2]^{3/2}}$

For more practice, please see: Study Questions: SQ1-3 (p. 139-141)
Practice Questions: PQ1-14 (p. 144-146)

Study Question – Chapter 11 (Equilibrium)

7A. For the reaction: $2AB(g) \rightleftharpoons A_2(g) + B_2(g)$,
the reaction vessel shown is at equilibrium. What is the equilibrium partial pressure of AB(g) when starting with 1.0 atm of AB(g) and no product at this temperature?



(A) 0.16 **(B)** 0.20
(C) 0.40 **(D)** 0.42

Work:

$K_p = \frac{P_{A_2} \cdot P_{B_2}}{P_{AB}^2}$ from the pic $= \frac{4 \cdot 4}{2^2} = 4$

R	$2AB \rightarrow$	$A_2 +$	B_2
I	1.00	0	0
C	$-2x$	x	x
E	$1.00 - 2x$	x	x

$4 = \frac{x^2}{(1.00-2x)^2}$ if we take the $\sqrt{\ }$ of the whole thing:

Could also guess & check

$\sqrt{4} = 2 = \frac{x}{1.00-2x}$ Solve for X

$2(1.00-2x) = x$

$2.00 - 4x = x$ $2.00 = 5x$ $x = 2/5 = \boxed{0.40}$

Practice Question – Chapter 11 (Equilibrium)

7B. For the reaction of: $CH_4(g) + 2O_2(g) \rightleftharpoons CO_2(g) + 2H_2O(g)$ +heat $\Delta H < 0$,
which change will increase $[CO_2]$ **because** K increases?

(A) increase $[CH_4]$ **(B)** decrease $[CH_4]$ **(C)** increase T **(D)** decrease T

Work:

ΔH negative so exothermic take heat away will increase $[CO_2]$ which is the one that changes K

For more practice, please see: Study Questions: SQ4-7 (p. 141-144)
Practice Questions: PQ15-30 (p. 146-149)

Study Question – Chapter 12 (Acids and Bases)

8A. In this equation: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$, what is HCO_3^- and what is its conjugate?

(A) HCO_3^- is an acid, and its conjugate base is H_2CO_3.

(B) HCO_3^- is an acid, and its conjugate base is OH^-.

[✓ **(C)**] HCO_3^- is a base, and its conjugate acid is H_2CO_3.

(D) HCO_3^- is a base, and its conjugate acid is OH^-.

Work:

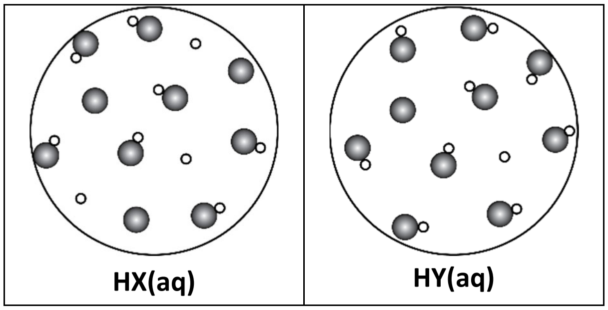
$$\underbrace{HCO_3^-}(aq) + H_2O(l) \rightleftharpoons \underbrace{H_2CO_3}(aq) + OH^-(aq),$$

Differ by a proton

HCO_3^- is a conjugate base to H_2CO_3
proton acceptor!

Practice Question – Chapter 12 (Acids and Bases)

8B. You have two acids, 1.0 M HX and 1.0 M HY, which are shown in the particulate diagrams to the right. What are the relative values of K_a and pH?



(A) $K_a(HX) > K_a(HY)$ $pH(HX) > pH(HY)$

(B) $K_a(HX) < K_a(HY)$ $pH(HX) > pH(HY)$

[✓ **(C)**] $K_a(HX) > K_a(HY)$ $pH(HX) < pH(HY)$

(D) $K_a(HX) < K_a(HY)$ $pH(HX) < pH(HY)$

Work:

10 total

HX is: 7 associated 3 dossociated

HY: 9 associated ; only 1 dissociated

HX is a stronger acid than HY

$Ka(HX) > Ka(HY)$ $pH(Hx) < pH(HY)$

For more practice, please see: Study Questions: SQ1, 3 (p. 151-153)
Practice Questions: PQ3-9, 13-15 (p. 156-158)

Study Question – Chapter 12 (Acids and Bases)

9A. What is the pH of 0.145 M CH_3COOH? $K_a(CH_3COOH) = 1.8\times10^{-5}$

(A) 1.95 **(B)** 2.80 **(C)** 3.91 **(D)** 5.58

Work:

R	$CH_3COOH + H_2O_{(l)} \rightleftharpoons$	H_3O^+ +	CH_3COO^-
I	0.145	0	0
C	$-x$	x	x
E	$0.145-x$	x	x

$$1.8\times10^{-5} = \frac{x^2}{0.145} \qquad x = 1.62\times10^{-3}$$

$$\boxed{pH = 2.79}$$

Practice Question – Chapter 12 (Acids and Bases)

9B. What is the pH of 0.575 M ethyl amine ($CH_3CH_2NH_2$)? $K_b(CH_3CH_2NH_2) = 5.6\times10^{-4}$

(A) 1.746 **(B)** 3.492 **(C)** 10.508 **(D)** 12.254

sig figs issue!

Work:

R	$B + H_2O \rightleftharpoons$	OH^- +	BH^+
I	0.575	0	0
C	$-x$	x	x
E	$0.575-x$	x	x

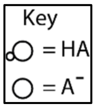
$$5.6\times10^{-4} = \frac{x^2}{0.575} = 1.79\times10^{-2}$$

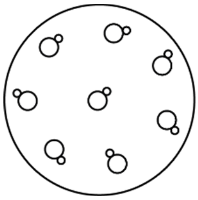
$$pOH = 1.75 \quad = \quad pH = 12.25$$

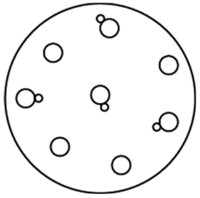
For more practice, please see: Study Questions: SQ2 (p. 152-153)
Practice Questions: PQ9-12, 18 (p. 157-158)

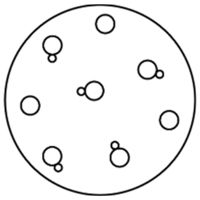
Study Question – Chapter 12 (Acids and Bases)

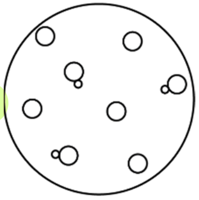
10A. Which buffer has a pH greater than the pK_a? *Water molecules have been omitted for clarity.*



(A)


(B)


(C)


(D)


Work:

more acid so < pka

= so pH = pka

5 > 2 so < pka

3 : 5 so > pka

Practice Question – Chapter 12 (Acids and Bases)

10B. What is the pH of a buffer made by combining 50.0 mL of 1.00 M C_6H_5COOH and 5.35 g of NaC_6H_5COO?

pK_a(C_6H_5COOH)	4.19
MM(NaC_6H_5COO)	132.1 g·mol^{-1}

(A) 4.10 **(B)** 4.28 **(C)** 5.00 **(D)** 5.42

Work:

$pH = pka + \log B/A$

$\frac{5.35 g}{} \Big| \frac{1 mol}{132.1 g} \Big| \frac{}{0.050 L} = 0.810\ M$ base

$pH = 4.19 + \log \frac{0.810}{1.00} = \boxed{4.10}$

For more practice, please see: Study Questions: SQ7 (p. 155-156)
Practice Questions: PQ21-25 (p. 158-159)

Study Question – Chapter 12 (Acids and Bases)

11A. For the titration of 1.0 M benzoic acid (C_6H_5COOH, $K_a = 6.5\times10^{-5}$) with 1.0 M NaOH, what is the pH at the equivalence point?

(A) 5.06 **(B)** 8.94 **(C)** 9.09 **(D)** 9.81

Work:

$\frac{1.0 \text{ mol}}{2x \text{ volume}} = 0.5 M$ conjugate base

$K_b = K_w/K_a = 1.54\times10^{-10}$

R	$B + H_2O \rightleftharpoons$	OH^-	$+ BH^+$
I	0.50	0	0
C	$-x$	x	x
E	$0.50-x$	x	x

$1.54\times10^{-10} = \frac{x^2}{0.50}$

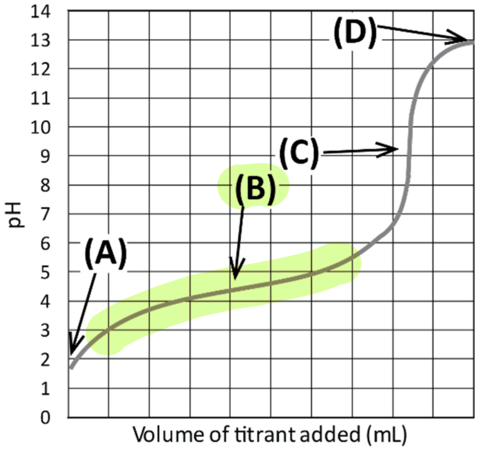
$x = 8.77\times10^{-6}$

pOH = 5.06

pH = 14 − pOH = $\boxed{8.94}$

Practice Question – Chapter 12 (Acids and Bases)

11B. On the titration curve shown, where is the buffer region?



Work:

For more practice, please see: Study Questions: SQ8 (p. 156)
Practice Questions: PQ26-30 (p. 159-160)

Study Question – Chapter 12 (Acids and Bases)

12A. Which solution has a pH greater than 7?

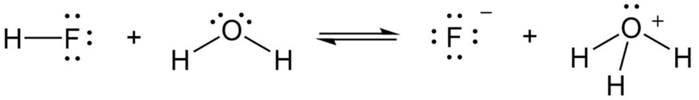
(A) $AlCl_3$(aq)
(B) $NaNO_3$(aq)
(C) NH_4Cl(aq)
[✓ **(D)** $NaNO_2$(aq)]

Work:

HNO_2 is a weak acid, therefore its conjugate NO_2^- is a weak base pH > 7

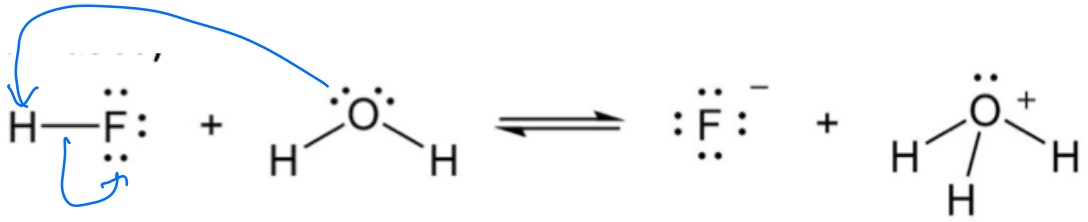
Practice Question – Chapter 12 (Acids and Bases)

12A. For the reaction of HF and H_2O (shown), what is the Lewis acid and why?



(A) HF because it is an electron pair donor.
[✓ **(B)** HF because it is an electron pair acceptor.]
(C) H_2O because it is an electron pair donor.
(D) H_2O because it is an electron pair acceptor.

Work:



Oxygen acts as the lone pair donor on the H_2O, so the lewis base
HF accepts the lone pair, thus would be the Lewis acid

For more practice, please see: Study Questions: SQ6 (p. 154-155)
Practice Questions: PQ1-2, 19-20 (p. 156-158)

Study Question – Chapter 13 (Solubility Equilibria)

13A. In which solution would $[Ag^+]$ be the lowest? $K_{sp}(AgCl) = 1.6\times10^{-10}$ $K_{sp}(AgBr) = 7.7\times10^{-13}$

(A) 1.0 g AgCl in 100.0 mL of water

(B) 1.0 g AgCl in 100.0 mL of 1 M NaCl

(C) 1.0 g AgBr in 100.0 mL of water

[✓ **(D)**] 1.0 g AgBr in 100.0 mL of 1 M NaBr

Work:

Ksp is lower for AgBr so $[Ag^+]$ would be less

in the presence of the common ion Br^- from NaBr, the equilibrium would shift further towards

reactants

Practice Question – Chapter 13 (Solubility Equilibria)

13B. If you combine equal volumes of 2.0×10^{-6} M $Ni(NO_3)_2$ and 2.0×10^{-6} M of NaOH, what is the value of Q_{sp} and what will happen? $K_{sp}(Ni(OH_2)) = 6.5\times10^{-18}$ Should be $Ni(OH)_2$

[✓ **(A)**] $Q_{sp} = 1\times10^{-18}$ and, because $Q_{sp} < K_{sp}$, the solution is unsaturated.

(B) $Q_{sp} = 1\times10^{-18}$ and, because $Q_{sp} < K_{sp}$, the solution is saturated.

(C) $Q_{sp} = 8\times10^{-18}$ and, because $Q_{sp} > K_{sp}$, the solution is unsaturated.

(D) $Q_{sp} = 8\times10^{-18}$ and, because $Q_{sp} > K_{sp}$, the solution is saturated.

Work:

$Ni^{2+}_{(aq)} + 2OH^- \longrightarrow Ni(OH)_{2(s)}$ equal volumes = [ ]original/2

$Q_{sp} = [Ni^{2+}][OH^-]^2 = 1.0\times10^{-6} \cdot (1.0\times10^{-6})^2 = 1.0\times10^{-18}$

$Q_{sp} < K_{sp}$ so solution is unsaturated

For more practice, please see: Study Questions: SQ1-8 (p. 163-169)
Practice Questions: PQ1-30 (p. 169-172)

Study Question – Chapter 14 (Thermodynamics)

14A. What is the change in entropy for the formation of $NO_2(g)$:

$$\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g)$$

***S*° / J·mol⁻¹·K⁻¹**	
$N_2(g)$	191.61
$O_2(g)$	205.14
$NO_2(g)$	240.06

(A) $-156.69\ J{\cdot}K^{-1}$ **(B)** $-60.89\ J{\cdot}K^{-1}$ **(C)** $226.53\ J{\cdot}K^{-1}$ **(D)** $636.81\ J{\cdot}K^{-1}$

Work:

$\Delta S^\circ_{rxn} = \Delta S^\circ_{prod} - \Delta S^\circ_{reactants}$

$= [1mol \cdot 240.06] - [\frac{1}{2}mol \cdot 191.61 + 1mol\ 205.14] = -60.9\ J/K$

Practice Question – Chapter 14 (Thermodynamics)

14B. Which reaction or process has $\Delta S > 0$?

(A) $H_2(g) \rightarrow 2H(g)$ $+\Delta S$

(B) $\overbrace{H_2(g) + \frac{1}{2}O_2(g)}^{\text{1.5 moles gas}} \rightarrow \overbrace{H_2O(g)}^{\text{1mol}}$ $-\Delta S$

(C) $CO_2(g) \rightarrow CO_2(aq)$ $-\Delta S$

(D) $H_2O(l) \rightarrow H_2O(s)$ $-\Delta S$

Work:

For more practice, please see: Study Questions: SQ1-2 (p. 176-177)
Practice Questions: PQ1-8 (p. 181-182)

Study Question – Chapter 14 (Thermodynamics)

15A. What is the standard molar free energy of combustion for pentane (C_3H_8(g))?

ΔG°_f / kJ·mol^{-1}	
C_3H_8(g)	–23.49
CO_2(g)	–394.36
H_2O(l)	–237.13

(A) –2155.09 kJ·mol^{-1} [✓ **(B)** –2108.11 kJ·mol^{-1}] **(C)** –654.98 kJ·mol^{-1} **(D)** –608.00 kJ·mol^{-1}

Work: $C_3H_8 + 5O_2 \longrightarrow 4H_2O + 3CO_2$ products are H_2O (l) + CO_2. You will need to balance it.

$\Delta G^\circ_{rxn} = [4 \cdot -237.13 + 3\,mol \cdot -394.36] - [1mol \cdot -23.49 + 5\,mol \cdot 0] = -2108\ kJ/mol$

Practice Question – Chapter 14 (Thermodynamics)

15B. The dimerization of NO_2 is exothermic: $2NO_2(g) \rightleftharpoons N_2O_4(g)$

What can be concluded about the spontaneity of this reaction?

$\Delta S^\circ_{rxn} = (-)$ more order/less random
$\Delta H^\circ_{rxn} = (-)$ exothermic

(A) It is spontaneous at all temperatures.

[✓ **(B)**] It is spontaneous at low temperatures but not at high temperatures.

(C) It is spontaneous at high temperatures but not at low temperatures.

(D) It is not spontaneous at any temperature.

Work:

$\Delta G^\circ = \Delta H - T\Delta S$
$= (-) - [+ \cdot -]$ So at high temps will get more positive or non-spon

For more practice, please see: Study Questions: SQ4-6, 8 (p. 177-180)
Practice Questions: PQ11-19, 25-28 (p. 182-184)

Study Question – Chapter 14 (Thermodynamics)

16A. What is the melting point of aluminum?

(A) 156 K [✓ **(B)** 941 K]

(C) 1063 K **(D)** 6428 K

	ΔH°_f / kJ·mol^{-1}	S° / J·K^{-1}·mol^{-1}
Al(s)	0	28.33
Al(l)	10.56	39.55

Work:

$Al_{(s)} \longrightarrow Al_{(l)}$

$\frac{\Delta H}{\Delta S} = T = \frac{10.56\ kJ/mol}{0.01122\ kJ/mol\ K} = \boxed{941\ K}$

$\Delta G = 0$ at the phase change!

$\Delta H = 10.56\ kJ/mol$

$\Delta S = 39.55 - 28.33 = 11.22\ J/mol\ K = 0.01122\ kJ/mol\ K$

Practice Question – Chapter 14 (Thermodynamics)

16B. For the decomposition of N_2O_4(g):

$N_2O_4(g) \rightarrow N_2(g) + 2O_2(g)$ $\Delta G^\circ_{rxn} = -97.89$ kJ

What are the relative values of K and ΔG at 25°C when the partial pressures of gases are as shown in the table?

N_2O_4 = 1 atm
N_2 = 0.01 atm
O_2 = 0.01 atm

(A) $K < 1$ and $|\Delta G| > |\Delta G^\circ_{rxn}|$ **(B)** $K < 1$ and $|\Delta G| < |\Delta G^\circ_{rxn}|$

[✓ **(C)** $K > 1$ and $|\Delta G| > |\Delta G^\circ_{rxn}|$] **(D)** $K > 1$ and $|\Delta G| < |\Delta G^\circ_{rxn}|$

Work: $\Delta G^\circ = -RT \ln K$

$\frac{-97.89\ kJ}{-8.314 \times 10^{-3}\ kJ/mol\ K \cdot 298\ K} = 39.5 = \ln K$ $K = e$^ $39.5 = 1.4 \times 10^{17}$

$K > 1$

* Conceptual =

mostly reactants so should be more spontaneous so ΔG = more $(-)$

$\Delta G = \Delta G^\circ + RT \ln Q$

$= -97.89 + 8.314 \times 10^{-3} \cdot 298 \cdot \ln\left(\frac{0.01^2 \cdot 0.01}{1}\right) = -132\ kJ$

$\Delta G > \Delta G^\circ$

For more practice, please see: Study Questions: SQ3, 7, 9 (p. 177, 179-181)
Practice Questions: PQ9-10, 20-24, 29-30 (p. 182-184)

Study Question – Chapter 17 (Electrochemistry)

17A. What is the coefficient for water when the reaction is balanced in basic solution?

$ClO^-(aq) + Cr(s) \rightarrow Cl^-(aq) + Cr^{3+}(aq)$ ***unbalanced***

(A) 1 **(B)** 2 [✓ **(C)** 3] **(D)** 6

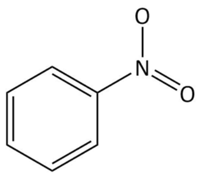
Work:

$(2e^- + \underbrace{2H^+ + ClO^-}_{+1} \rightarrow \underbrace{Cl^- + H_2O}_{-1}) \times 3$

$(Cr_{(s)} \rightarrow Cr^{3+} + 3e^-) \times 2$

Practice Question – Chapter 17 (Electrochemistry)

17B. What is the oxidation state of nitrogen in nitrobenzene (shown to the right)?



☆

(A) +5 [✓ **(B)** +3]

(C) +2 **(D)** +1

Work:

N more EN than C, so gives N −1 charge

However, O more EN than N, so takes charge from N so

NO_2 would be −1

so

N: 1x_ = +3

O: 2x−2 = −4

overall −1

For more practice, please see: Study Questions: SQ1 (p. 188)
Practice Questions: PQ1-3 (p. 193)

Study Question – Chapter 17 (Electrochemistry)

18A. What combination will yield the greatest E°_{cell}?

E°_{red} / V			
$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$	+0.34	$In^{3+}(aq) + 3e^{-} \rightarrow In(s)$	–0.34
$Cu^{2+}(aq) + e^{-} \rightarrow Cu^{+}(aq)$	+0.19	$Cr^{2+}(aq) + 2e^{-} \rightarrow Cr(s)$	–0.91

(A) $3Cu^{2+}(aq) + In^{3+}(aq) \rightarrow 3Cu^{+}(aq) + In(s)$

(B) $3Cu^{2+}(aq) + 2In(s) \rightarrow 3Cu(s) + 2In^{3+}(aq)$

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

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

Work:

$$E_{cell} = E_{cathode} - E_{anode}$$

$$= 0.34 - -0.91$$

$$\uparrow_{red} \qquad \uparrow_{oxid}$$

$$Cu^{2+} + 2e^{-} \rightarrow Cu(s)$$

$$Cr(s) \rightarrow Cr^{2+} + 2e^{-}$$

Practice Question – Chapter 17 (Electrochemistry)

18B. What is the strongest reducing agent?

E°_{red} / V			
$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$	+0.34	$Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s)$	–0.74

(A) $Cu^{2+}(aq)$

(B) $Cr^{3+}(aq)$

(C) $Cu(s)$

(D) $Cr(s)$

Work:

The more positive (+) the reduction potential, E°_{red}, the greater the tendency to be reduced. A reducing agent is a substance that causes the reduction of something else. So Cr(s) is the best choice here because Cu^{2+} is the most easily reduced from the list.

the most easy to reduce (Cu^{2+} here) has the product which is hardest to oxidize (not the stronger reducing agent)

For more practice, please see: Study Questions: SQ2-3 (p. 189)
Practice Questions: PQ4-11 (p. 193-194)

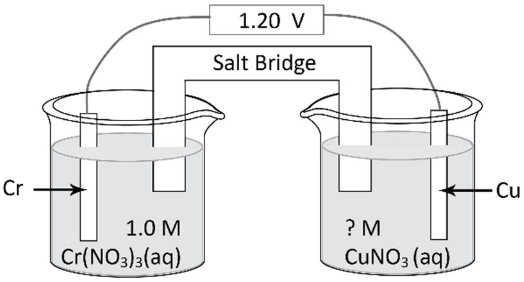
Study Question – Chapter 17 (Electrochemistry)

19A. What is the value of Q to yield a measured cell potential of E_{cell} = 1.20 V?

$$3Cu^{+}(aq) + Cr(s) \rightarrow Cr^{3+}(aq) + 3Cu(s)$$

$+3e^-$

E^{o}_{cell} = 1.26 V



(A) 0.00091

(B) 0.097

(C) 10.3

(D) 1100

Work:

$$E_{cell} = E^{o}_{cell} - \frac{0.0592}{n} \log Q$$

$$1.20V = 1.26V - \frac{0.0592}{3} \log Q$$

$$1.20 - 1.26 \times \frac{3}{0.0592} = +3.04 = \log Q$$

$$Q = 10^{\wedge +3.04} = 1.1 \times 10^{3} = 1100$$

Practice Question – Chapter 17 (Electrochemistry)

19A. What mass of tin can be deposited from a $Sn(NO_3)_2$ solution when applying a current of 1.05 A for 95 s? The molar mass of tin is 118.7 g·mol^{-1}.

(A) 0.061 g

(B) 0.12 g

(C) 0.25 g

(D) 1.3 g

Work:

$$Sn^{2+} + 2e^- \rightarrow Sn^{o}(s)$$

$$1.05 A \cdot 95 s = 9.975 \times 10^{1} A \cdot s = C$$

Faradys constant = 9.648×10^{4} C/mole$^-$

$$mole^- = \frac{1 mole^-}{9.649 \times 10^{4}} \times 9.975 \times 10^{1} C = 1.03 \times 10^{-3} mole^-$$

$$\frac{1.03 \times 10^{-3} mole^-}{} \bigg| \frac{1 mol\ Sn^{2+} reduced}{2 mole^-} \bigg| \frac{118.71 g}{1 mol\ Sn} = \boxed{0.061 g\ Sn}$$

For more practice, please see: Study Questions: SQ5, 7-8 (p. 191-193)
Practice Questions: PQ14-20, 24-30 (p. 195-196)

Study Question – Chapter 16 (Nuclear Chemistry)

20A. What is missing from the nuclear reaction shown? $^{253}_{99}\text{Es} + ^{4}_{2}\alpha \rightarrow \underline{\ ?\ } + ^{1}_{0}\text{n}$

(A) $^{249}_{97}\text{Bk}$ **(B)** $^{250}_{97}\text{Bk}$ [✓ **(C)** $^{256}_{101}\text{Md}$] **(D)** $^{257}_{101}\text{Md}$

Work: $^{A}_{Z}\mathcal{E}$ notation = $^{\text{mass \#}}_{\text{proton \#}}\text{Element}$

$^{253}_{99}\text{Es} + \frac{4}{2}\alpha = ^{256}_{101}\text{Md} + ^{1}_{0}\text{n}$

top # & bottom # just have to add up

Practice Question – Chapter 16 (Nuclear Chemistry)

20B. When radium-226 decays to form radon-222, what particle is emitted?

[✓ **(A)** $^{4}_{2}\alpha$] **(B)** $^{1}_{1}\text{p}$ **(C)** $^{1}_{0}\text{n}$ **(D)** $^{0}_{-1}\beta$

Work:

$^{226}_{88}\text{Ra} \rightarrow ^{222}_{86}\text{Rn} + ^{4}_{2}\text{He}$ = alpha particle $^{4}_{2}\alpha$

For more practice, please see: Study Questions: SQ1-3 (p. 199-201)
Practice Questions: PQ1-10 (p. 201-202)

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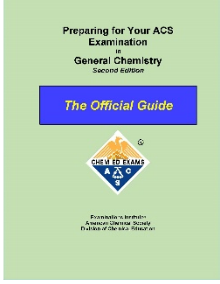

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