ENTER YOUR ANSWER TO EACH QUESTION ON THE ANSWER SHEET, IN
PENCIL. THERE IS NO ADDITIONAL PENALTY FOR INCORRECT ANSWERS.

1.   Your version of this test is **VERSION A**. You need to identify that you are
       answering this version of the test by filling in the bubble for answer (A) for
       question 1 on the optical scan sheet. Your answer for question 1 is:

(A)

Questions 2 through 16 are each worth 2 marks.

2.   For the following reaction at 25°C

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$
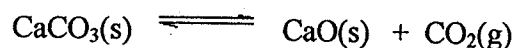
       Find the **FALSE** statements:

i. $K_p = K_c$ for this reaction.
ii.   Increasing the total pressure on the system will not change the value of K_p.
iii.   If the concentration of CO_2 is doubled, the equilibrium constant will be halved.
iv.   Equilibrium is reached when the rates of forward and reverse reactions are equal.
v. $K_c = [CO_2][CaO]$

(A)   i, iii, v

(B)   i, ii, iv

(C)   i, iii, iv, v

(D)   ii, iii, iv

(E)   ii, iii, v

$K_p = P_{CO_2}$

$K_c = [CO_2]$

$P_{CO_2} = [CO_2] RT$

1

0.5

3. At a certain temperature, the equilibrium constant, K_c, for the reaction

$$SO_3(g) + NO(g) \rightleftharpoons NO_2(g) + SO_2(g)$$

is found to be 0.500. If 0.300 moles each of SO_3 and NO are placed in 2.00 L flask and allowed to react at that temperature, the equilibrium concentration of SO_2, in mol L^{-1}, will be

(A) 0.0620

(B) 0.0500

(C) 0.124

(D) 0.100

(E) 0.0310

Handwritten work:

$$SO_3(g) + NO(g) \rightleftharpoons NO_2(g) + SO_2(g)$$

INITIAL CONC. $\dfrac{0.300\,mol}{2\,L}$ $\quad$ $\dfrac{0.300\,mol}{2\,L}$ $\quad$ — $\quad$ —

change $\quad -X \quad\quad -X \quad\quad X \quad\quad X$

equil conc. $\dfrac{0.300-X}{2}$ $\quad$ $\dfrac{0.300-X}{2}$ $\quad$ $\dfrac{X}{2}$ $\quad$ $\dfrac{X}{2}$

$$0.500 = \frac{\left(\frac{X}{2}\right)^2}{\left(\frac{0.300-X}{2}\right)^2} = \left(\frac{X}{0.300-X}\right)^2$$

$$X = 0.124 \, mol \qquad \frac{X}{2} = 0.0620$$

4. Consider the equilibrium :

$$CCl_4(g) \rightleftharpoons C(s) + 2Cl_2(g)$$
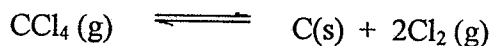
At 800 K pure CCl_4 was placed in a closed container with an initial pressure of CCl_4 equal to 2.00 atm. After equilibrium has been reached, the total pressure is 2.45 atm. What is K_p, in atm, for the equilibrium at 800 K?

(A) 0.367

(B) 0.818

(C) 0.0343

(D) 0.523

(E) 0.581

Handwritten work:

$$CCl_4(g) \rightleftharpoons C(s) + 2Cl_2(g)$$

INITIAL $2.00\,atm$ $\quad\quad\quad +2X$

change $-X$ $\quad\quad\quad\quad 2X$

equil. $2.00-X$

$$2X + 2.00 X = 2.45 \, atm.$$
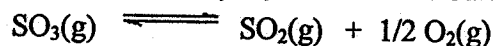
$$X = 0.45 \, atm$$

$$K_p = \frac{P_{Cl_2}^2}{P_{CCl_4}} = \frac{(2X)^2}{2.00-X} = \underline{0.523 \, atm}$$

5. For the following reaction:

$$2 SO_3(g) \rightleftharpoons 2 SO_2(g) + O_2(g)$$

the equilibrium constant, $K_c = 1.6 \times 10^{-10}$. Calculate the value of K_c for

$$SO_3(g) \rightleftharpoons SO_2(g) + 1/2\ O_2(g)$$

(A) 8.0×10^4

(B) 1.0

(C) 1.3×10^{-5}

(D) 1.6×10^{-6}

(E) 6.25×10^9

$$1.6 \times 10^{-10} = \frac{[O_2][SO_2]^2}{[SO_3]^2} = K_{C_1}$$

$$K_{C_2} = \frac{[O_2]^{\frac{1}{2}}[SO_2]}{[SO_3]}$$

$$K_{C_2} = \sqrt{K_{C_1}}$$

6. For the following reaction, $K_p = 1.1 \times 10^{-3}$ at 1000 K.

$$3 H_2(g) + Fe_2O_3(s) \rightleftharpoons 2 Fe(s) + 3 H_2O(g)$$
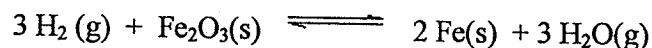
Which of the following statements are **FALSE**?

T i.   Adding $H_2(g)$ to the equilibrium, at constant temperature, causes the partial pressure of $H_2O(g)$ to increase.

T ii. Adding $Fe_2O_3(s)$ to the equilibrium mixture, at constant temperature, does not change the partial pressure of hydrogen gas.

F iii. $K_p = 19.1 \times 10^2$ is the equilibrium constant for the reverse reaction at the same temperature.

F iv. Doubling the volume of an equilibrium mixture at constant temperature causes the partial pressure of $H_2O(g)$ to increase.

T v.   The only variable that can change the value of equilibrium constant is temperature.

0.5

(A) ii, iii

(B) iii, iv

(C) iv, v

0.5

(D) i, v

(E) i, ii

$$K_{p_1} = \frac{P_{H_2O}^3}{P_{H_2}^3}$$

K_p for reverse reaction:

$$K_{p_2} = \frac{P_{H_2}^3}{P_{H_2O}^3} = 909$$

7. Consider the following system, which has already reached equilibrium:

$$NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$$

Which of the following changes will cause the equilibrium to shift to the **RIGHT**?

F (i)   Adding more $NH_3(g)$.
F (ii)   Adding $NH_4HS(s)$.
T (iii)   Increasing the volume of the container.
F (iv)   Adding a catalyst.
T (v)   Removal of $NH_3(g)$ from the system by reaction with HCl(aq).

0.5 ↑

(A)   (ii) and (iii)

(B)   (iii) and (v)
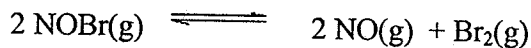
(C)   (i) and (v)

(D)   (ii) and (iv)

0.5 ↓ (E)   (i) and (iv)

i) equilibrium will shift to the left

ii) will not change equilibrium (solid)

iv) " " " position of equilibrium

$$K_p = P_{NH_3} \times P_{H_2S}$$

8. NOBr(g), initially at pressure P_o, dissociates according to the equilibrium:

$$2\,NOBr(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$$

When equilibrium is established at 25°C, NOBr(g) is 34% dissociated and the equilibrium total pressure, P, is 0.25 atm. K_p, in atm, for this equilibrium is
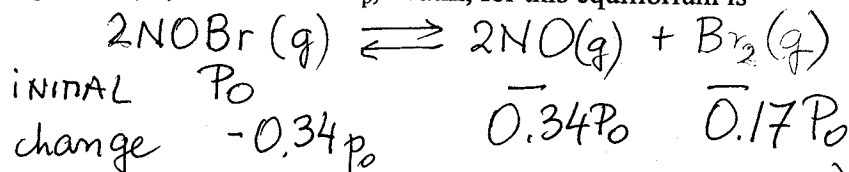
(A)   0.088

(B)   0.21

(C) 9.8×10^{-6}

(D)   0.019

(E)   0.0096

$$2\,NOBr(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$$

INITIAL P_o
change $-0.34\,P_o$ $0.34\,P_o$ $0.17\,P_o$

TOTAL PRESSURE $= 0.25\,atm = (P_o - 0.34\,P_o) + 0.34\,P_o + 0.17\,P_o$

$$0.25\,atm = 1.17\,P_o$$

$$P_o = 0.213\,atm.$$

$$K_p = \frac{P_{NO}^2 \times P_{Br_2}}{(P_{NOBr})^2} = 9.6 \times 10^{-3}$$

9. In a certain experiment, 50.0 mL of HBr react completely with 0.086 mol of Na_2O to produce sodium bromide and water. What was the original concentration of the HBr?

(A) 3.40×10^{-3} M

(B) 4.30×10^{-3} M

(C) 0.860 M

(D) 1.72 M

(E) 3.44 M

balanced equation:

$$2\,HBr_{(aq)} + Na_2O_{(s)} \rightarrow 2NaBr_{(aq)} + H_2O_{(l)}$$
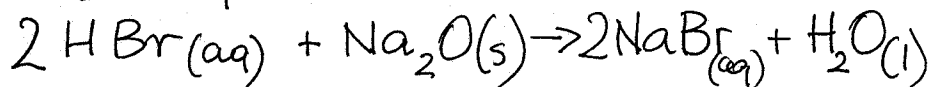
0.086 mol Na_2O reacts with 0.172 moles HBr

$$M_{HBr} = \frac{0.172\,mol}{0.0500\,L} = 3.44$$

10. The order of **decreasing acidity** of the oxo-acids of chlorine which decreases from left to right is: (strongest acid → weakest acid).

(pg. 621 in CHANG)

(A) $HClO_4 > HClO_3 > HClO_2 > HClO$

(B) $HClO > HClO_2 > HClO_3 > HClO_4$

(C) $HClO > HClO_3 > HClO_4 > HClO_2$

(D) $HClO_2 > HClO > HClO_3 > HClO_4$

(E) $HClO_3 > HClO_2 > HClO > HClO_4$

11. A 50.0 mL solution of 0.0500 M NaOH was treated with 700.0 mL of 0.00750 M HNO_3 and diluted to 2.00 L with water. What is the pH of the 2.00 L solution?

1.5 (A) 2.46

(B) 2.56

(C) 2.86

(D) 6.76

(E) 11.1

moles $NaOH = 2.5 \times 10^{-3}$

moles $HNO_3 = 5.25 \times 10^{-3}$

equation: $NaOH_{(aq)} + HNO_{3(aq)} = NaNO_{3(aq)} + H_2O_{(l)}$

INITIAL 2.5×10^{-3} 5.25×10^{-3}

2.75×10^{-3} moles HNO_3 left

$[HNO_3] = \dfrac{2.75 \times 10^{-3} \text{ moles}}{2L} = 1.38 \times 10^{-3} M$

12. Carbon dioxide dissolves in water and reacts to give HCO_3^- in the equilibrium shown below. Blood transports CO_2 to the lungs and has a pH = 7.4. What is the **ratio** of $[HCO_3^-]$ to $[CO_2]$ in blood at pH 7.4?

$$CO_2 + 2 H_2O \rightleftharpoons HCO_3^- + H_3O^+ \quad K_a = 4.3 \times 10^{-7}$$

(A) 10.8

(B) 4.3

(C) 1.7

(D) 0.093

1.5 (E) 5.8×10^{-8}

pH = 7.4 $[H_3O^+] = 3.98 \times 10^{-8}$

$4.3 \times 10^{-7} = \dfrac{[HCO_3^-] \times 3.98 \times 10^{-8}}{[CO_2]}$

$\dfrac{[HCO_3^-]}{[CO_2]} = 10.8$

13. A 0.300 M solution of hydrazine hydrochloride, $N_2H_5^+ Cl^-$, had a pH of 4.50.
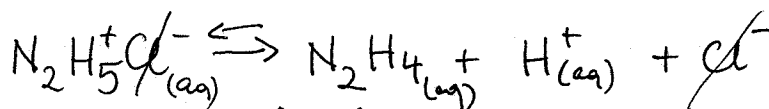Calculate the K_b for hydrazine, N_2H_4.

(A) 1.05×10^{-4}

(B) 3.20×10^{-5}

(C) 3.00×10^{-6}
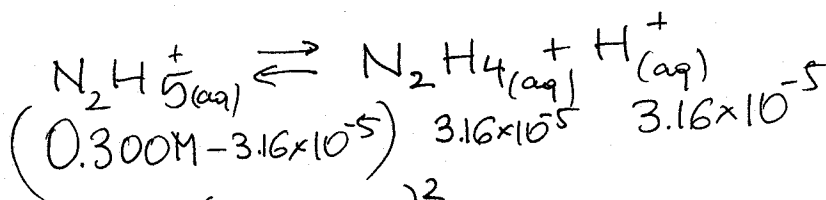
(D) 3.33×10^{-9}

1.5 (E) 3.13×10^{-10}

$$N_2H_5^+ Cl^-_{(aq)} \rightleftharpoons N_2H_{4(aq)} + H^+_{(aq)} + Cl^-$$

$$pH = 4.50 \therefore [H^+] = 3.16 \times 10^{-5} M$$

$$N_2H_5^+_{(aq)} \rightleftharpoons N_2H_{4(aq)} + H^+_{(aq)}$$

$$(0.300M - 3.16 \times 10^{-5}) \quad 3.16 \times 10^{-5} \quad 3.16 \times 10^{-5}$$

$$K_a = \frac{(3.16 \times 10^{-5})^2}{(0.300 - 3.16 \times 10^{-5})} = 3.33 \times 10^{-9}$$

$$K_a \times K_b = K_{H_2O} \qquad K_b = 3.00 \times 10^{-6}$$

14. A 50.00 mL sample of a weak acid HA was titrated to the equivalence point with
29.40 mL of 0.300 M NaOH. K_a for HA $= 2.9 \times 10^{-6}$. Calculate the pH at the
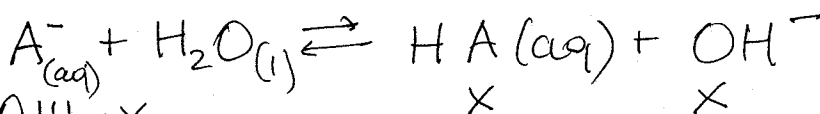equivalence point.

1 (A) 4.63

(B) 4.71

(C) 9.29

(D) 9.51

(E) 10.8

$$HA_{(aq)} + NaOH_{(aq)} \rightarrow Na^+_{(aq)} + A^-_{(aq)} + H_2O_{(l)}$$

$$\# \text{ moles NaOH} = 8.82 \times 10^{-3}$$

$$\text{Volume} = 50.00mL + 29.40mL = 79.40mL$$

$$M_{(A^-)} = \frac{8.82 \times 10^{-3} mol}{0.07940 L} = 0.111 M$$

$$A^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HA_{(aq)} + OH^-$$

$$\text{equil. } 0.111 - x \qquad\qquad x \qquad\quad x$$

$$\frac{x^2}{0.111 - x} = 3.44 \times 10^{-9}$$

$$x = [OH^-] = 1.95 \times 10^{-5}$$

$$pOH = 4.71 \qquad pH = 9.29$$

15. "Concentrated" ammonia is a saturated solution of gaseous ammonia in water with a molarity of 17.0. What is the percentage ionization of ammonia in this solution? $K_b = 1.8 \times 10^{-5}$ for ammonia.
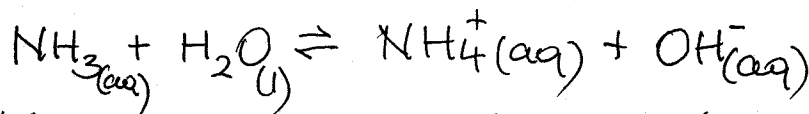
1.5

↑

(A)   0.0010

(B)   0.0018

(C)   0.018

(D)   0.10

(E)   0.18

$$1.8 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

17 - x   x   x

$$\frac{x^2}{17-x} = 1.8 \times 10^{-5} \qquad x^2 = 3.06 \times 10^{-4}$$

$$x = 0.017_{(5)}$$

$$\% \ ionization = \frac{[NH_4^+]}{INITIAL\ CONCENTR.} \times 100\% = \frac{0.017_{(5)}}{17} \times 100 = 0.1C$$

16. What is the pH of the solution at the half equivalence point of the titration of 50.0 mL of a 0.010 M solution of potassium propanoate with 0.010 M HBr (after ½ the required number of moles is added)? The K_a for propanoic acid (conjugate acid of potassium propanoate) is 1.3×10^{-5}.

(A)   2.00

(B)   2.79

(C)   3.44

(D)   4.89

(E)   9.12

↳ 1

at half equivalence point

$$[HA] = [A^-]$$

(buffer solution)

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

[A⁻] → [conjugate base]

[HA] → [acid]

$$pH = pK_a$$

Questions 17 through 21 are worth 3 marks each.

17. In a certain experiment, 0.750 atm H_2 and 0.340 atm I_2 are placed in a 5.00 L flask at 375 °C, and allowed to come to equilibrium according to:

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

K_p for the reaction is 1.64×10^{-3}. How many **moles of HI** are present at equilibrium?

(A) 2.00×10^{-2} mol

(B) 1.88×10^{-3} mol

(C) 2.63×10^{-3} mol

(D) 3.25×10^{-3} mol

(E) 3.76×10^{-3} mol

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

initial P	0.750 atm	0.340 atm	—
change	$-x$	$-x$	$+2x$
equil.	$0.750 - x$	$0.340 - x$	$2x$

$$K_p = \frac{(2x)^2}{(0.750 - x)(0.340 - x)} = 1.64 \times 10^{-3}$$

$$4x^2 = 1.64 \times 10^{-3}\,(0.255 - 1.09x + x^2)$$

$$x = 0.0100 \text{ atm}$$

$$P\,HI \text{ (at equilibrium)} = 0.0200 \text{ atm}$$

$$\#moles\ HI = \frac{PV}{RT} = 1.88 \times 10^{-3} mol$$

18. This summer while at home you discover an old can in the basement that contains 50.0 g of lye (NaOH(s)). Rather than throw it in the garbage, to be "green", you decide to first neutralize the lye and pour the resulting solution down the drain.

How many litres of vinegar (5% CH_3COOH by weight) would you need to treat the lye to reach the equivalence point? Assume no volume changes on addition of the solid lye. The K_a for acetic acid is 1.8×10^{-5}. Assume that density of CH_3COOH is 1.00 g mL^{-1}.

(A) 0.75 L

(B) 1.0 L

1 ← (C) 1.3 L
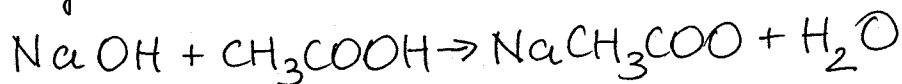
(D) 1.5 L

(E) 3.0 L

50g NaOH = 1.25 moles

NaOH + CH_3COOH → $NaCH_3COO$ + H_2O

1.25 moles CH_3COOH needed

5% CH_3COOH = 5g CH_3COOH in 100g of solution

density CH_3COOH = 1g mL^{-1}

5% CH_3COOH → 5g CH_3COOH in 100 mL solution

Mmass CH_3COOH = 60 g mol^{-1}

∴ 0.833 mol/L

Molarity = $\dfrac{\#\ moles}{Volume}$

0.833 M = $\dfrac{1.25}{V}$

V = 1.5 L

19. A solution of 0.01 M NH_3 is titrated with 0.01 M HCl. Which of the indicators below would you use? The K_b for NH_3 is 1.8×10^{-5}.

Answer	Indicator	pH range	Colour Change	pK_a
i	Methyl orange	3.3 - 4.6	red - yellow	4.2
ii	Methyl red	4.2 - 6.2	red - yellow	5.2
iii	Bromthymol blue	6.0-7.8	Yellow - blue	7.2
iv	Thymol blue	7.9 - 9.4	Yellow - blue	8.2
v	Phenolphthalein	8.3 - 10.0	Colourless - red	9.5

(A) i only

(B) ii only

(C) iii only

1.5 (D) ii or iv

(E) v only

at equiv. point $0.005M$ $NH_4^+{}_{(aq)}$

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_3{}_{(aq)} + H_3O^+$

$0.005-x \qquad\qquad x \qquad x$
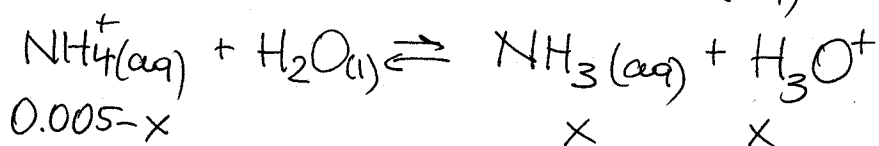
$\dfrac{x^2}{0.005-x} = \dfrac{1\times10^{-14}}{1.8\times10^{-5}}$

$x = [H_3O^+] = 1.66 \times 10^{-6}$

$\underline{pH = 5.8}$ only Methyl red changes pH

20. What is the approximate pH of the endpoint of the titration of 100.0 mL of 0.050M HNO_3 by 0.050 M KOH if phenolphthalein (see previous question) is used as an indicator. The pK_a for phenolphthalein is given in the table in Question 19, above.

(A) 13

(B) 10

(C) 8.3

2.5 (D) 7.0

(E) There is insufficient information provided to answer this question.

titration: strong acid vs strong base, pH = 7
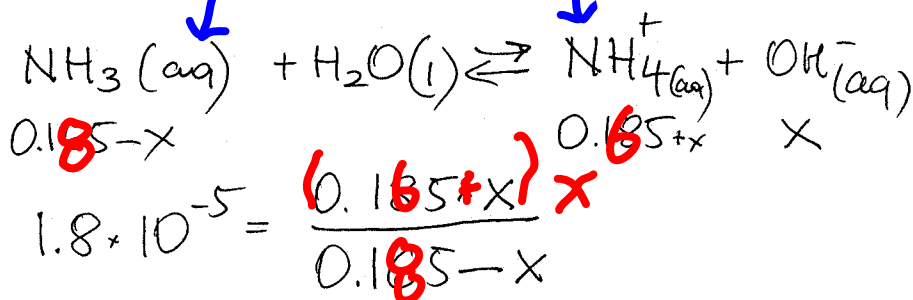but at 8.3 pH phenolphtalein is still colourless. It is red at pH = 10.

21. What is the pH of the solution resulting from the addition of 7.80 grams of solid KOH to 750 mL of 0.350 M NH_4Br? Assume no volume change occurs on addition of KOH solid. The K_b for NH_3 is 1.8×10^{-5}.

1.5

(A) 4.79

(B) 4.85

(C) 9.30

(D) 11.5

(E) 13.3

7.80g KOH → 0.139 moles KOH

$$NH_4^+ (aq) + OH^- (aq) \longrightarrow NH_3 (aq) + H_2O (l)$$

INITIAL 0.750L × 0.350M   ↓

0.263 moles   0.139 moles

FINAL (0.263 − 0.139) — 0.139 moles

0.123 moles

FINAL concentr. 0.165 M NH_4^+ 0.185 M NH_3

$$NH_3 (aq) + H_2O (l) \rightleftharpoons NH_4^+ (aq) + OH^- (aq)$$

0.185 − x   0.185 + x   x

$$1.8 \times 10^{-5} = \frac{(0.165 + x)\, x}{0.185 - x}$$

$[OH^-] = 2.02 \times 10^{-5}$

$pOH = 4.69$

$pH = 9.30_5$