Name: KEY

Chemistry 30
Equilibrium, Acids and Bases Workbook

Chemical Equilibrium

1. For each of the following, write the chemical reaction equation with the appropriate equilibrium arrows.
   a) pH measurements indicate that acetic acid in vinegar is approximately 1% ionized into hydrogen ions and acetate ions.

   $$CH_3COOH(aq) \xrightleftharpoons{\sim 1\%} H^+(aq) + CH_3COO^{1-}(aq)$$

   b) Quantitative analysis of the reaction of sodium sulphate and calcium chloride solutions shows that the products are favoured.

   $$Na_2SO_4(aq) + CaCl_2(aq) \xrightleftharpoons{>50\%} NaCl(aq) + CaSO_4(s)$$

   c) Aluminum sulphate solution reacts quantitatively with a sodium hydroxide solution.

   $$Al_2(SO_4)_3(aq) + NaOH(aq) \longrightarrow Al(OH)_3(s) + Na_2SO_4(aq)$$

   d) An analysis of the reaction between sodium hypochlorite and ammonium nitrate shows that the reactants are favoured.

   $$NaOCl(aq) + NH_4NO_3(aq) \xrightleftharpoons{<50\%} NaNO_3(aq) + NH_4OCl(aq)$$

2. Write the equilibrium law for each of the following chemical reaction equations.
   a) $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$ $\quad K_c = \dfrac{[SO_3(g)]^2}{[SO_2(g)]^2 \cdot [O_2(g)]}$

   b) $2\ NO_2(g) \rightleftharpoons 2\ NO(g) + O_2(g)$ $\quad K_c = \dfrac{[NO(g)] \cdot [O_2(g)]}{[NO_2(g)]^2}$

   c) $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$ $\quad K_c = \dfrac{[NH_3(g)]^2}{[N_2(g)] \cdot [H_2(g)]^3}$

   d) $\cancel{Na_2}SO_4(aq) + Ca\cancel{Cl_2}(aq) \rightleftharpoons 2\ \cancel{NaCl}(aq) + CaSO_4(s)$

   $$Ca^{2+}(aq) + SO_4^{2-}(aq) \rightleftharpoons CaSO_4(s) \qquad K_c = \frac{1}{[Ca^{2+}(aq)] \cdot [SO_4^{2-}(aq)]}$$

3. Chlorine and carbon monoxide gases are mixed in a 1.00 L container. Initially, 1.5 mol of chlorine was present with excess carbon monoxide. The percent reaction is 53%.
   a) Write the chemical reaction equation including the appropriate equilibrium arrows and the percent reaction.

$$Cl_2(g) + CO(g) \xrightleftharpoons{53\%} \underset{\text{(from Part c)}}{COCl_2(g)}$$

   b) Write the equilibrium law for this reaction.

$$K_c = \frac{[COCl_2(g)]}{[Cl_2(g)][CO(g)]}$$

   c) At equilibrium, 0.80 mol of $COCl_2(g)$, 1.75 mol of carbon monoxide and 0.70 mol of chlorine were present. Calculate the equilibrium constant.

$$K_c = \frac{[0.80]}{[0.70][1.75]} \qquad K_c = 0.65$$

4. In an experiment at 200°C, 0.500 mol/L of hydrogen bromide gas is placed in a sealed container and it decomposes into hydrogen gas and bromine gas.
   a) Write the equilibrium equation and law for this reaction.

$$2\,HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$$

$$K_c = \frac{[H_2(g)]\cdot[Br_2(g)]}{[HBr(g)]^2}$$

   b) The equilibrium concentrations in this system are: $[HBr(g)] = 0.240$ mol/L, $[H_2(g)] = [Br_2(g)] = 0.130$ mol/L. Calculate the equilibrium constant.

$$K_c = \frac{(0.130)(0.130)}{(0.240)^2}$$

$$\boxed{= 0.293}$$

Graphical Analysis

The Haber-Bosch process of the industrial production of ammonia involves the equilibrium

$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g) + 92.2\text{ kJ}.$$

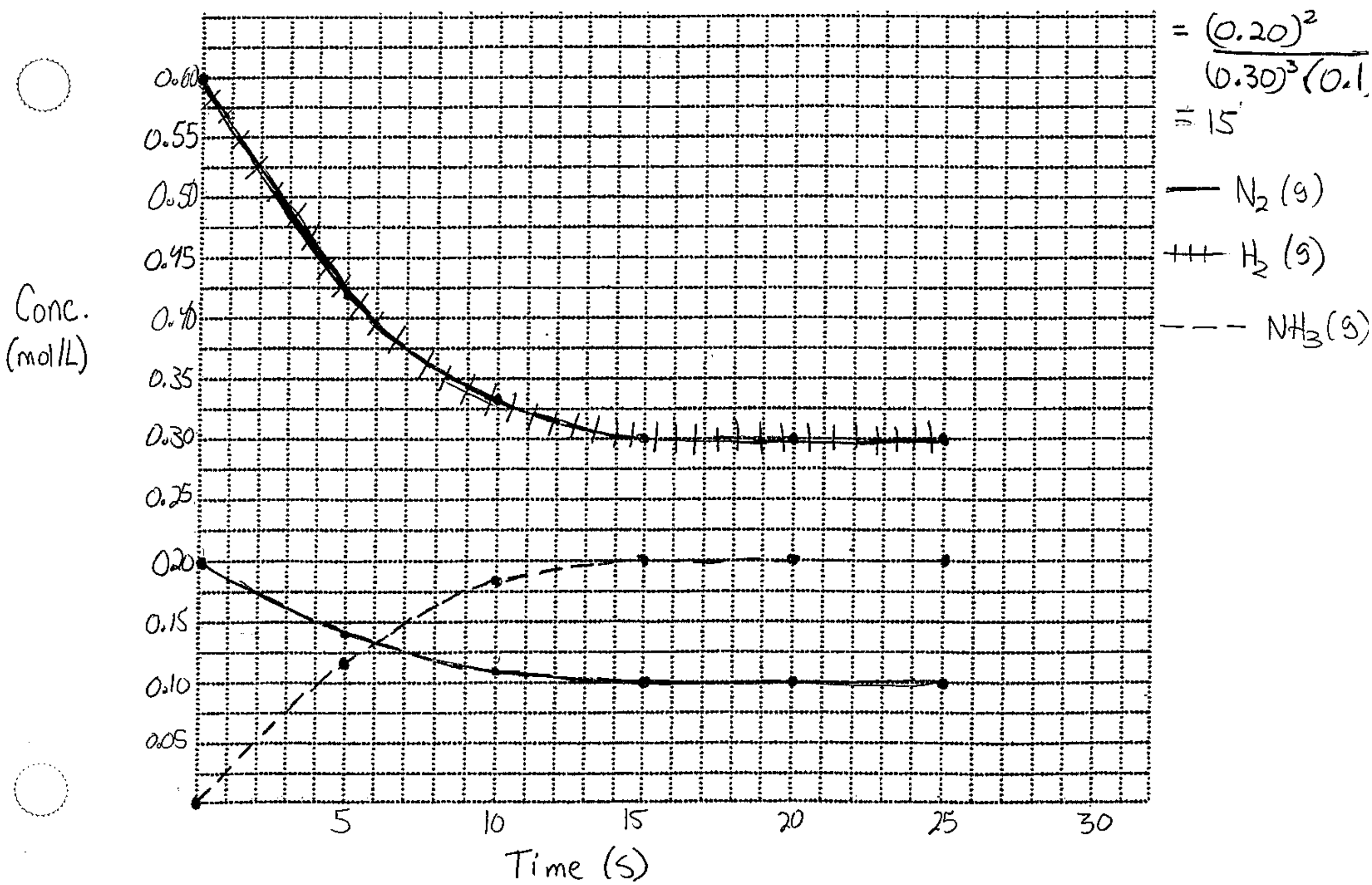
In a laboratory experiment designed to study this equilibrium, a chemical engineer injects 0.20 mol of $N_2(g)$ and 0.60 mol of $H_2(g)$ into a 1.0 L flask at 500°C. She records her analysis of the flask at 5 s intervals in the table shown.

Time (s)	**Concentration (mol/L)**		
	$N_2(g)$	**$H_2(g)$**	**$NH_3(g)$**
0	0.20	0.60	0.00
5	0.14	0.42	0.12
10	0.11	0.33	0.18
15	0.10	0.30	0.20
20	0.10	0.30	0.20
25	0.10	0.30	0.20

Analyze the data by:

1. Draw a graph of the concentrations of $N_2(g)$, $H_2(g)$ and $NH_3(g)$ versus time on the graph paper below. Include a legend with your graph.
2. State the time required for equilibrium to be established 15 s
3. **Calculate** the equilibrium constant for this reaction…showing all work.

$$K_c = \frac{[NH_3(g)]^2}{[H_2(g)]^3 \cdot [N_2(g)]} = \frac{(0.20)^2}{(0.30)^3(0.1} \doteq 15$$



Le Châtelier's Principle

1. Nitrogen monoxide, a major air pollutant, is formed in automobile engines from the endothermic reaction of nitrogen gas and oxygen gas.
   a) Write the equilibrium reaction equation including the term "energy" in the equation.

   $$N_2(g) + O_2(g) + \text{energy} \rightleftharpoons 2NO(g)$$

   b) What is the direction of the equilibrium shift if the concentration of oxygen is increased? – to products

   c) What is the direction of the equilibrium shift if the pressure is increased?

   – no effect

   d) What is the direction of the equilibrium shift if the temperature is decreased?

   – to reactants

   e) Gasoline burns better at higher temperatures. What is one disadvantage of operating automobile engines at higher temperatures?

   – more heat shifts equilibrium to products ∴ more $NO(g)$ pollution

2. In a sealed container, nitrogen monoxide gas and oxygen gas react to form nitrogen dioxide gas and are allowed to come to equilibrium. The reaction of nitrogen monoxide and oxygen is exothermic.
   a) Write the equilibrium reaction equation including the term "energy" in the equation.

   $$\underbrace{2NO(g) + O_2(g)}_{3\text{ moles gas}} \rightleftharpoons \underbrace{2NO_2(g)}_{2\text{ moles gas}} + \text{energy}$$

   b) What is the direction of the equilibrium shift if the temperature is decreased.

   – towards products

   c) What is the direction of the equilibrium shift if the $[NO_{(g)}]$ is decreased.

   – towards reactants

   d) What is the direction of the equilibrium shift if the $[NO_{2(g)}]$ is increased.

   – towards reactants

   e) What is the direction of the equilibrium shift if the <u>volume of the system</u> is <u>decreased</u>.

   – towards products (pressure increased)

3. The following is an equilibrium mixture:

   $$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

   faint yellow colourless red

   Predict the colour change in the mixture when each of the following changes is made:
   a) a crystal of KSCN(s) is added to the system – towards products
      ↳ high solubility ∴ adding $K^{+}(aq)$ & $\underline{SCN^{1-}(aq)}$
   b) a crystal of $FeCl_3(s)$ is added to the system – towards products
      $\underline{Fe^{3+}}(aq)$ & $Cl^{1-}(aq)$
   c) a crystal of NaOH(s) is added to the system
      $OH^{1-}(aq)$ & $Fe^{3+}(aq) \rightarrow Fe(OH)_3(s)$ – towards reactants
      ∴ $Fe^{3+}(aq)$ decreases

Le Châtelier's Principle

—from Hebden: Chemistry 12

1. Use Le Châtelier's Principle to describe the effect of the following changes on the position of the equilibrium:

   A. The equilibrium is: $N_2O_3(g) \rightleftharpoons NO(g) + NO_2(g)$

      a) increase the [NO] - to reactants

      b) increase the $[N_2O_3]$ - to products

      c) add a catalyst - no effect

      d) increase the pressure by <u>decreasing the volume</u> - to reactants

   B. The equilibrium is: $\underset{4\text{ mol}}{2\,H_2(g) + 2\,NO(g)} \rightleftharpoons \underset{3\text{ mol}}{N_2(g) + 2\,H_2O(g)}$

      a) decrease the $[N_2]$ - to products

      b) decrease the pressure by <u>increasing the volume</u> - to reactants

      c) decrease the [NO] - to reactants

   C. The equilibrium is: $\underset{3\text{ mol}}{2\,CO(g) + O_2(g)} \rightleftharpoons \underset{2\text{ mol}}{2\,CO_2(g)} + 566\text{ kJ}$

      a) increase the temperature - to reactants

      b) add a catalyst - no effect

      c) increase the $[O_2]$ - to products

   D. The equilibrium is: $\underset{2\text{ mol}}{I_2(g) + Cl_2(g)} \rightleftharpoons \underset{2\text{ mol}}{2\,ICl(g)}\ \Delta H = +35.0\text{ kJ}$

      + energy

      a) decrease the temperature - to reactants

      b) decrease the $[Cl_2]$ - to reactants

      c) increase the pressure by <u>decreasing the volume</u> - no effect

2. Describe the effect (increase, decrease or no change) on the concentration of the bold-faced substance by each of the given changes.

   A. The equilibrium is: $\underset{4\text{ mol}}{N_2(g) + \mathbf{3\,H_2(g)}} \rightleftharpoons \underset{3\text{ mol}}{2\,NH_3(g)}$ $\Delta H = -92.0\text{ kJ}$

      a) increase the $[N_2]$ - decrease

      b) increase the volume - increase

      c) increase the temperature - increase

      d) add a catalyst - no change

B. The equilibrium is: $\overset{2\,mol}{2\,HF(g)}$ $\rightleftharpoons$ $\overset{2\,mol}{\mathbf{F_2(g)}}$ + $H_2(g)$ $\Delta H = +536.0$ kJ

+energy

a) decrease the [H_2] – increase
b) decrease the volume – no effect
c) decrease the temperature – decrease

C. The equilibrium is: $SnO_2(s) + 2\,CO(g)$ $\rightleftharpoons$ $Sn(s) + \mathbf{2\,CO_2(g)}$ $\Delta H = +13.0$ kJ

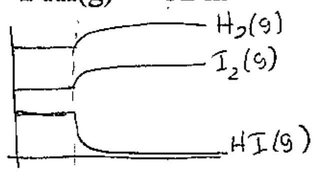
+ energy

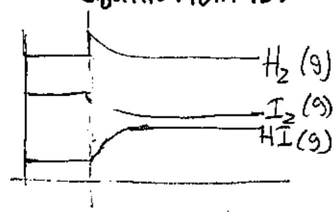
a) increase the [CO] – increase
b) add a catalyst – no effect
c) increase the temperature – increase

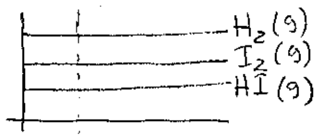
3. Show the equilibrium adjustment in each of the following situations graphically:
***Note that the relative positioning of the molecules is not relevant *What happens to each to restore equilibrium is.

A. The equilibrium is: $H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g) + 52$ kJ

a) increase the temperature
b) decrease the volume
c) inject some $H_2(g)$
d) add a catalyst

a)


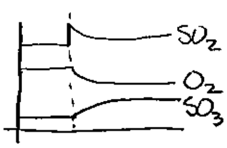
c)


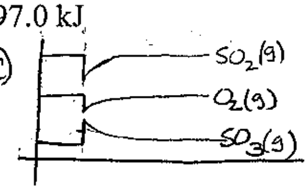
b)


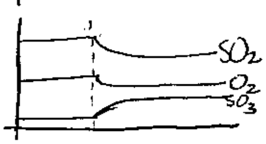
d) same as b)

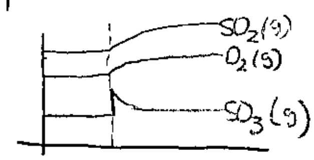
B. The equilibrium is: $2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$ $\Delta H = -197.0$ kJ

a) inject some $SO_2(g)$
b) decrease the temperature
c) increase the volume
d) increase the [$SO_3(g)$]

a)


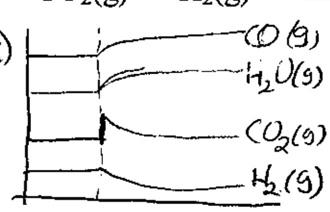
c)


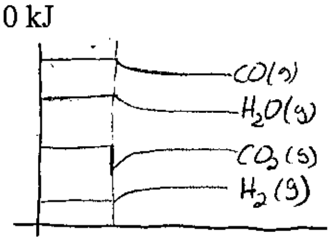
b)


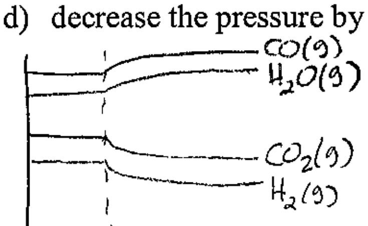
d)


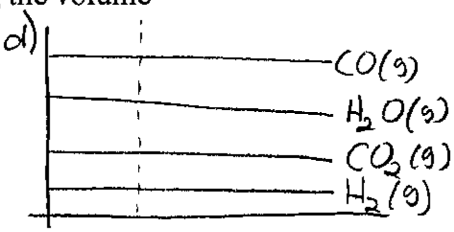
C. The equilibrium is: $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$ $\Delta H = -41.0$ kJ

a) inject some $CO_2(g)$
b) remove some $CO_2(g)$
c) increase the temperature
d) decrease the pressure by increasing the volume

a)


b)


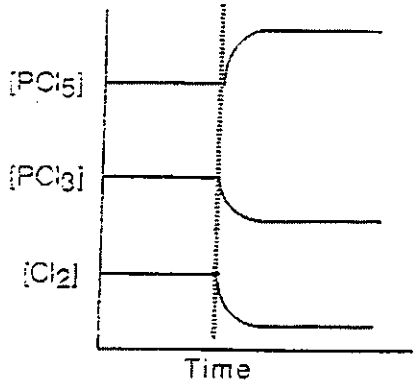
c)


d)


4. Interpret the following graphs in terms of the changes which must have been imposed on the equilibrium:

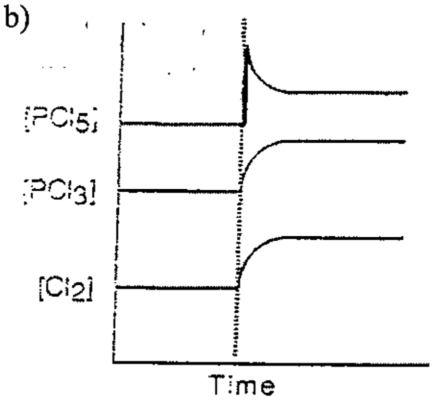
A. The equilibrium is: $PCl_5(g) + 92.5\ kJ \rightleftharpoons PCl_3(g) + Cl_2(g)$

a)



- temp. decreased

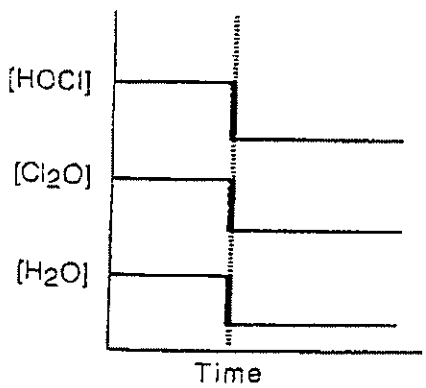
b)



- increase conc. of $PCl_5(g)$
($PCl_5(g)$ add)

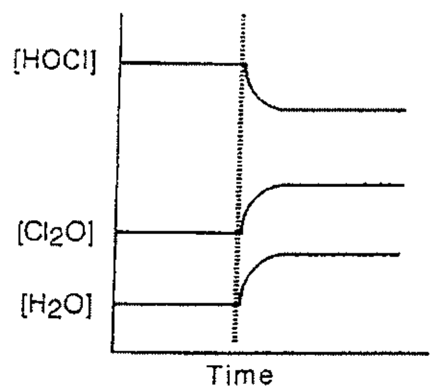
B. The equilibrium is: $H_2O(g) + Cl_2O(g) \rightleftharpoons 2\ HOCl(g) + 70\ kJ$

a)



- increase volume - decrease pressure

b)



- increase temperature

ICE Tables

1. 1.00 mol of hydrogen gas and 1.00 mol of iodine gas are sealed in a 1.00 L reaction vessel and allowed to react at 450°C. At equilibrium, 1.56 mol of hydrogen iodide gas is present. Calculate K_c for the reaction.

	$[H_2(g)]$	$[I_2(g)]$	$[HI(g)]$
I	1	1	0
C	-0.78	-0.78	+1.56
E	0.22	0.22	1.56

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

$K_c = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]} = \frac{(1.56)^2}{(0.22)(0.22)} = 50.3$

2. In an experiment, 2.00 mol of $H_2(g)$ and 2.00 mol of $F_2(g)$ are introduced into a 1.00 L flask at 500°C. After equilibrium was reached, the concentration of HF(g) was 0.500 mol/L. Calculate the K_c for this reaction at 500°C.

$H_2(g) + F_2(g) \rightleftharpoons 2HF(g)$

	$H_2(g)$	$F_2(g)$	$[HF(g)]$
I	2.00	2.00	0
C	-0.25	-0.25	+0.500
E	1.75	1.75	0.500

$K_c = \frac{[HF(g)]^2}{[H_2(g)] \cdot [F_2(g)]} = \frac{(0.50)^2}{(1.75)(1.75)} = 0.0816$

3. Phosphorus pentachloride gas decomposes into phosphorus trichloride gas and chlorine gas. If the $[PCl_5(g)]_i = 8.1 \times 10^{-3}$ mol/L and the $[PCl_3(g)]_i = 0.298$ mol/L, calculate the K_c. The $[Cl_2(g)]_{eq} = 2.00 \times 10^{-3}$ mol/L.
$K_c = 1.32 \times 10^{-5}$

$\boxed{\text{OMIT}}$

4. A 1.0 L flask was filled with 2.0 mol of $SO_2(g)$ and 2.0 mol of $NO_2(g)$ and heated to 250°C. If the K_c is 2.75, calculate the equilibrium concentrations of all species at 250°C.

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

	$[SO_2(g)]$	$[NO_2(g)]$	$[SO_3(g)]$	$[NO(g)]$
I	2	2	0	0
C	$-x$	$-x$	$+x$	$+x$
E	$2-x$	$2-x$	x	x

$2.75 = \frac{x^2}{2-x^2}$ $\sqrt{2.75} = \frac{x}{2-x}$ $1.658 = \frac{x}{2-x}$

$3.3166 - 1.658x = x$
$3.3166 = 2.658x$
$x = 1.25$

Equilibrium conc.
$SO_2 = 0.75$ mol/L
$NO_2 = 0.75$ mol/L
$SO_3 = 1.25$ mol/L
$NO = 1.25$ mol/L

5. At a certain temperature, $K_c = 2.6 \times 10^{-4}$ for the reaction of H_2O with Cl_2O. Calculate the equilibrium concentrations of each species if 10.0 g of H_2O and 20.0 g of Cl_2O are mixed in a sealed 2.00 L flask. $n = 0.555$ mol $n = 0.230$

$$H_2O(g) + Cl_2O(g) \rightleftharpoons 2\,HOCl(g)$$

Requires quadratic equation because not a perfect square:
We don't need to do these!

pH and pOH

1. Use K_W to calculate the $[H_3O^+(aq)]$ in a solution which has a $[OH^-(aq)]$ of 1.0×10^{-11} mol/L. (1.0×10^{-3} mol/L)

$$[H_3O^+(aq)] = \frac{K_W}{[OH^-(aq)]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-11}}$$
$$= \boxed{0.0010 \text{ mol/L}}$$

2. Use K_W to calculate the $[OH^-(aq)]$ in a solution which has a $[H^+(aq)]$ of 1.0×10^{-4} mol/L. (1.0×10^{-10} mol/L)

$$[OH^-(aq)] = \frac{K_W}{[H_3O^+(aq)]}$$
$$= \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-4}}$$
$$\boxed{= 1.0 \times 10^{-10} \text{ mol/L}}$$

3. Determine the pH from the following [$H_3O^+(aq)$] or [$OH^-(aq)$].
   a) [$H_3O^+(aq)$] = 1.0×10^{-13} mol/L pH = ________ acidic or basic ________
   b) [$H_3O^+(aq)$] = 1×10^{-2} mol/L pH = ________ acidic or basic ________
   c) [$H_3O^+(aq)$] = 10^{-6} mol/L pH = ________ acidic or basic ________
   d) [$OH^-(aq)$] = 0.0010 mol/L pH = ________ acidic or basic ________
   e) [$OH^-(aq)$] = 1.0×10^{-6} mol/L pH = ________ acidic or basic ________

See below

* 3a. 13 base 3b. 2 acid 3c. 6 acid 3d. 11 base 3e. 8 base Answers

4. Find the pH of a lime that has [$H_3O^+(aq)$] = 0.0120 mol/L. (1.921)

$$pH = -\log(0.0120)$$
$$= 1.921$$

5. Determine the pH of a blood sample with a [$OH^-(aq)$] = 2.6×10^{-7} mol/L. (7.41)

$$pOH = -\log(2.6 \times 10^{-7})$$
$$= 6.585$$
$$pH = 14 - 6.858 = 7.41$$

6. An ammonia solution has a pOH of 2.92. What is the concentration of hydroxide ions in solution? (0.00120 mol/L)

$$[OH^{1-}(aq)] = 10^{-pOH}$$
$$= 10^{-2.92}$$
$$= 0.00120 \text{ mol/L}$$

7. Find the [$H_3O^+(aq)$] and pH of a solution made by dissolving 10.0 g of KOH in water to make 4.00 L of solution. (2.24×10^{-13} mol/L, 12.649)

$$\frac{56.11g}{1mol} = \frac{10g}{x\,mol}$$
$$x = 0.17822$$

$$KOH(aq) \longrightarrow K^{1+}(aq) + OH^{1-}(aq)$$

$$\frac{0.17822mol}{4L} = 0.04455$$

0.04455 mol/L

$$[H_3O^+(aq)] = \frac{1.00 \times 10^{-1}}{0.04455} = 2.24 \times 10^{-13} \text{ mol}$$

$$pH = -\log(2.24 \times 10^{-13}) = 12.649$$

8. A sodium hydroxide solution is prepared by dissolving 2.50 g to make 2.00 L of solution. Calculate both the hydroxide ion and hydrogen ion concentrations. (OH = 0.03125 mol/L, H = 3.2×10^{-13} mol/L)

$$\frac{40.00g}{1mol} = \frac{2.50g}{x\,mol}$$
$$x = 0.0625mol$$

$$NaOH(aq) \rightarrow Na^{1+}(aq) + OH^{1-}(aq)$$

$$\frac{0.0625mol}{2L} = 0.03125 \text{ mol/L}$$

0.03125 mol/L

$$[H_3O^+(aq)] = \frac{1.00 \times 10^{-14}}{0.03125} = 3.20 \times 10^{-13} \text{ mol/}$$

9. A 0.728 g sample of hydrogen chloride gas is dissolved in 200 mL of water. Calculate both the hydroxide ion and hydrogen ion concentrations. (H= 0.0998mol/L, OH = 1.002 x 10^{-13} mol/L)

$HCl(aq) \rightarrow H^{+}(aq) + OH^{1-}(aq)$

$\frac{36.46\text{ g}}{1\text{ mol}} = \frac{0.728\text{ g}}{x\text{ mol}}$

$x = 0.01996\ 7\text{ mol}$

$\frac{0.019967\text{ mol}}{0.200\text{ L}} = 0.0998\text{ mol/L}$

0.0998 mol/L 0.0998 mol/L

$[OH^{1-}(aq)] = \frac{1.00 \times 10^{-14}}{0.0998} = 1.002 \times 10^{-13}\text{ mol/L}$

$\boxed{1.00 \times 10^{-13}\text{ mol/L}}$

10. Calculate the pH of a solution made by dissolving 7.50 g of strontium hydroxide to make 500 mL of solution. (13.392) $Sr(OH)_2(aq) \longrightarrow Sr^{2+}(aq) + 2OH^{1-}(aq)$

$\frac{121.64\text{ g}}{1\text{ mol}} = \frac{7.50\text{ g}}{x\text{ mol}}$

$x = 0.06166\text{ mol}$

$\frac{0.0616657\text{ mol}}{0.500\text{ L}}$

$C = 0.1233\text{ mol/L}$

0.2466 mol/L

$pOH = -\log(0.2466) = 0.607955$

$pH = 14 - 0.607955 = 13.392$

11. Complete the following table:

$[H_3O^{+}(aq)]$	$[OH^{-}(aq)]$	pH	pOH	Acid/Base/Neutral
4.0 x 10^{-6} mol/L	2.5×10^{-9}	5.40	8.60	acid
3.2×10^{-10}	3.2×10^{-5}	9.50	4.50	base
5.0×10^{-4}	2.0 × 10^{-11} mol/L	3.30	10.70	acid
10 mol/L	1.0×10^{-15}	-1.00	15.00	acid
2.74×10^{-6}	3.66×10^{-9}	5.563	8.437	acid
2.3×10^{-13}	0.044	12.64	1.36	basic

Brønsted-Lowry Acids and Bases

1. Label each reactant in the following equations as Brønsted-Lowry acid or base:

a) $HSO_3^{-}(aq) + H_2O(l) \rightleftharpoons H_3O^{+}(aq) + SO_3^{2-}(aq)$

acid base C. acid C. base

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

base acid C. acid C. base

c) $HF(aq) + HSO_3^{-}(aq) \rightleftharpoons F^{-}(aq) + H_2SO_3(aq)$

acid base C. base C. acid

d) $H_2SO_3(aq) + HS^{-}(l) \rightleftharpoons HSO_3^{-}(aq) + H_2S(aq)$

acid base C. base C. acid

2. For each of the following reactions, label each reactant as Brønsted-Lowry acid or base **and** complete the reaction.

a) HNO_2 (aq) + Cl^-(l) ⇌ NO_2^{1-}(aq) + HCl(aq)

acid base c. base c. acid

b) CH_3COOH (aq) + SO_4^{2-}(aq) ⇌ CH_3COO^{1-}(aq) + HSO_4^{1-}(aq)

acid base c. base c. acid

c) H_2S (aq) + NO_2^-(aq) ⇌ HS^{1-}(aq) + HNO_2(aq)

acid base c. base c. acid

d) HCO_3^-(aq) + S^{2-} (l) ⇌ HS^{1-}(aq) + CO_3^{1-}(aq)

acid base c. acid c. base

e) CO_3^{2-}(aq) + H_2O(l) ⇌ HCO_3^{1-}(aq) + OH^{1-}(aq)

base acid c. acid c. base

3. For questions 1 and 2, label each product as conjugate acid or conjugate base.

Strengths of Acids

1. Calculate the pH of 500 mL of each of the following acids:

a) 1.0×10^{-2} mol/L HCl(aq)

strong acid ∴ $[H_3O^+(aq)] = 1.0 \times 10^{-2}$ mol/L } pH = 2.00

assumption rule!

$\frac{[HA]}{K_a}$

b) 6.00 mol/L HNO_2(aq)

✓ $\frac{6.00}{5.6 \times 10^{-4}} = 10714X$

$K_a = \frac{[H_3O^+(aq)]^2}{[HA]}$ $5.6 \times 10^{-4} = \frac{X^2}{6.00}$ $X = 0.057$

pH = 1.23

c) 1.50 mol/L H_2SO_3(aq)

$\frac{1.5}{1.4 \times 10^{-2}} = 107X$

Fails assumption rule ∴ don't have to solve. Requires ICE table & quadratic.

d) 6.8×10^{-2} mol/L HNO_3(aq)

strong acid ∴ $[H_3O^+(aq)] = 6.8 \times 10^{-2}$ mol/L } pH = 1.17

e) 6.3×10^{-1} mol/L HF(aq)

$\frac{6.3 \times 10^{-1}}{6.3 \times 10^{-4}} = 1000X$

$K_a = \frac{[H_3O^+(aq)]^2}{[HA(aq)]}$ $6.3 \times 10^{-4} = \frac{X^2}{6.3 \times 10^{-1}}$

$X = 0.0199$

pH = 1.70

2. A 0.80 mol/L solution of an unknown acid, HX(aq), has a pH of 3.75. $[H_3O^+(aq)] = 1.778 \times 10^{-4}$

a) Calculate the % reaction. $p = \frac{[H_3O^+(aq)]}{[HA(aq)]} \times 100 = \frac{1.778 \times 10^{-4}}{0.8} \times 100$

% reaction = 0.022%

b) Calculate the K_a, the acid ionization constant.

	HX	$H_3O^+(aq)$
I	0.80	0
C	-1.7782×10^{-4}	
E	0.7998	1.778×10^{-4}

$HX(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + X^{1-}(aq)$

$K_a = \frac{[H_3O^+(aq)]^2}{[HA(aq)]} = \frac{[1.7782 \times 10^{-4}]^2}{0.7998} = 4.0 \times 10^{-8}$

3. Calculate the pH of a solution containing 0.25 mol/L of an acid with an acid ionization constant (K_a) of 3.2×10^{-6} mol/L.

ASSUMPTION CHECK

$\frac{0.25}{3.2 \times 10^{-6}} = 78125$

$\therefore [HA(aq)]_{initial} = [HA(aq)]_{equilibrium}$

$K_a = \frac{[H_3O^+(aq)]^2}{[HA(aq)]}$ $\quad 3.2 \times 10^{-6} = \frac{x^2}{0.25}$

$x = 8.944 \times 10^{-4}$ mol/L

$pH = -\log(8.94427 \times 10^{-4})$

$= 3.05$

Strengths of Bases

1. Calculate the pH and pOH for each of the following:

a) 1.00 mol/L NaOH(aq)

1.00 mol/L $OH^{1-}(aq)$ pOH = 0.000 pH = 14.000

b) 1.00 mol/L $Ca(OH)_2(aq)$

2.00 mol/L $OH^{1-}(aq)$ pOH = −0.301 pH = 14.301

c) 0.650 mol/L $Al(OH)_3(aq)$

1.95 mol/L $OH^{1-}(aq)$ pOH = −0.290 pH = 14.290

d) A solution made by dissolving 5.82 g of barium hydroxide in 2.00 L of water.

$Ba(OH)_2 \rightarrow Ba^{2+}(aq) + 2OH^{1-}(aq)$

$\frac{171.35 g}{1 mol} = \frac{5.82 g}{x mol}$ $\quad x = 0.03396 / 2L = 0.01698$ mol/L $\quad$ 0.0339655 mol/L

pOH = 1.469

pH = 12.531

2. Calculate the K_b for each of the following bases at 25°C:

$K_b = \frac{K_w}{K_a}$

a) $NO_2^-(aq)$ $K_b = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-4}} = 1.8 \times 10^{-11}$

b) $F^-(aq)$ $K_b = \frac{1.00 \times 10^{-14}}{6.3 \times 10^{-4}} = 1.6 \times 10^{-11}$

c) $HSO_3^-(aq)$ $K_b = \frac{1.00 \times 10^{-14}}{1.4 \times 10^{-2}} = 7.1 \times 10^{-13}$

d) $HCO_3^-(aq)$ $K_b = \frac{1.00 \times 10^{-14}}{4.5 \times 10^{-7}} = 2.2 \times 10^{-8}$

e) $OOCCOO^{2-}(aq)$ $K_b = \frac{1.00 \times 10^{-14}}{1.5 \times 10^{-4}} = 6.7 \times 10^{-11}$

3. Calculate the pH of a 13.5 mol/L solution of $H_2PO_4^-(aq)$ using the following reaction:

$$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq)$$

assumption rule check

$\frac{13.5 \text{ mol/L}}{6.9 \times 10^{-3}} = 1956$ ✓

$\therefore [H_2PO_4^{1-}(aq)]$ initial $= [H_2PO_4^{1-}(aq)]$ equilibrium

Strengths of Acids and Bases – pH Calculations

Calculate the pH for each of the following solutions. Show all work.

1. 0.32 mol/L $Mg(OH)_2(aq)$ - strong base

$[OH^{1-}(aq)] = 2 \times 0.32 = 0.64$ mol/L

$pOH = -\log(0.64) = 0.19$

$pH = 14 - pOH = 14 - 0.1938 = \boxed{13.81}$

2. 6.00 mol/L $NH_3(aq)$

assumption rule $\frac{6.00}{1.7857 \times 10^{-7}}$ ✓

$K_b = \frac{[OH^{1-}(aq)]^2}{[WB]}$ $1.785 \times 10^{-5} = \frac{[OH^{1-}(aq)]^2}{6}$

$[OH^{1-}(aq)] = 0.010$

$[H_3O^+(aq)] = 9.6618 \times 10$

$pH = \boxed{12.015}$

3. 2.0×10^{-4} mol/L $KHSO_4(aq)$ **red with both litmus (acid)

assumpt check $\frac{2.0 \times 10^{-4}}{1.0 \times 10^{-2}}$

Ice table required → quadratic to solve ∴ don't need to do this.

4. 3.0 mol/L $H_2S(aq)$

assumption check $\frac{3.0}{8.9 \times 10^{-8}} = $ ✓

$K_a = \frac{[H_3O^+(aq)]^2}{[HA]}$ $8.9 \times 10^{-8} = \frac{x^2}{3}$ $x = 5.1672 \times 10^{-4}$ mol/L

$pH = -\log$

$= \boxed{3.29}$

5. 0.750 mol/L KF(aq) weak base F^{1-} $K_b = \frac{1.00 \times 10^{-14}}{6.3 \times 10^{-4}} = 1.587 \times 10^{-11}$

assumption check $\frac{0.750}{1.587 \times 10^{-11}} = $ ✓

$K_b = \frac{[OH^{1-}(aq)]^2}{[WB]}$ $1.587 \times 10^{-11} = \frac{[OH^{1-}(aq)]^2}{0.750}$ $[OH^{1-}(aq)] = 3.45 \times 10^{-6}$

$pOH = 5.46$

$\boxed{pH = 8.538}$

6. 0.0505 mol/L HI(aq)

Strong acid ∴ $[H_3O^+(aq)] = 0.0505$ mol/L

$pH = -\log(0.0505)$

$\boxed{pH = 1.297}$

7. 16 mol/L $CH_3COOH(aq)$

assumption check $\frac{16}{1.8 \times 10^{-5}}$ ✓

$K_a = \frac{[H_3O^+(aq)]^2}{[HA]}$ $1.8 \times 10^{-5} = \frac{[H_3O^+(aq)]^2}{16}$

$[H_3O^+(aq)] = 0.01697$

$\boxed{pH = 1.77}$

8. 2.00 mol/L NaCN(aq) $K_b = \frac{1.00 \times 10^{-14}}{6.2 \times 10^{-10}}$ $K_b = 1.6129 \times 10^{-5}$

assumption $\frac{2}{1.6129 \times 10^{-5}}$ ✓

$K_b = \frac{[OH^{1-}(aq)]^2}{[WB]}$ $1.6129 \times 10^{-5} = \frac{[OH^{1-}(aq)]^2}{2.00}$

$[OH^{1-}(aq)] = 0.0056$

$pOH = 2.245$

$\boxed{pH = 11.754}$

9. 5.7 mol/L $Na_2HPO_4(aq)$ **blue with both litmus $K_b = \frac{1.00 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6129 \times 10^{-7}$

assumption $\frac{5.7}{1.61 \times 10^{-7}}$ ✓

$K_b = \frac{[OH^{1-}(aq)]^2}{WB}$ $1.6129 \times 10^{-7} = \frac{[OH^{1-}(aq)]^2}{5.7}$ $[OH^{1-}(aq)] = 9.588 \times 10^{-4}$

$pOH = 3.01$

$\boxed{pH = 10.98}$

10. 0.750 mol/L $Al(OH)_3(aq)$

strong acid ∴ $[OH^{1-}(aq)] = 0.750 \times 3 = 2.25$ mol/L

$pOH = -0.352$

$\boxed{pH = 14.352}$

Predicting Acid-Base Reactions

For each of the following problems:

1. Write a reaction equation.
2. Label reactants as acid or base and products as conjugate acid or conjugate base.
3. Use appropriate arrow notation to indicate reaction predomination (forward or reverse).

1. Solutions of $Na_2SO_3(aq)$ and $HF(aq)$ are mixed in a beaker. $\underset{}{Na^{1+}(aq)}$, $\overset{SB}{SO_3^{2-}(aq)}$, $\overset{SA}{HF(aq)}$, $H_2O(l)$

$$\underset{A}{HF(aq)} + \underset{B}{SO_3^{2-}(aq)} \xrightleftharpoons{>50\%} \underset{CB}{F^{1-}(aq)} + \underset{CA}{HSO_3^{1-}}$$

2. A solution of $NH_4NO_3(aq)$ and a solution of $NaCH_3COO(aq)$ are mixed. $\underset{SA}{NH_4^{1+}(aq)}$, $\underset{SB}{CH_3COO^{1-}(aq)}$

$$\underset{A}{NH_4^{1+}(aq)} + \underset{B}{CH_3COO^{1-}(aq)} \xrightleftharpoons{<50\%} \underset{CB}{NH_3(aq)} + \underset{CA}{CH_3COOH(aq)}$$

3. Sodium benzoate is often used as a preservative. Write the equation for $NaC_6H_5COO(s)$ dissolving in a solution of $NaHSO_4(aq)$. $\underset{SB}{C_6H_5COO^{1-}(aq)}$, $\underset{SA}{HSO_4^{1-}(aq)}$

$$\underset{A}{HSO_4^{1-}(aq)} + \underset{B}{C_6H_5COO^{1-}(aq)} \xrightleftharpoons{>50\%} \underset{CB}{SO_4^{2-}(aq)} + \underset{CA}{C_6H_5COOH(aq)}$$

4. A household ammonia solution is mixed with a solution of nitrous acid. $\underset{SB}{NH_3(aq)}$, $\underset{SA}{HNO_2(aq)}$

$$\underset{A}{HNO_2(aq)} + \underset{B}{NH_3(aq)} \rightleftharpoons \underset{CB}{NO_2^{1-}(aq)} + \underset{CA}{NH_4^{1+}(aq)}$$

5. Nitric acid and potassium hydroxide solutions are mixed.

$H_3O^+(aq)$ $OH^{1-}(aq)$

$$\underset{A}{H_3O^+(aq)} + \underset{B}{OH^{1-}(aq)} \longrightarrow \underset{CA/B}{2H_2O(l)}$$

6. Sodium sulphate is dissolved into a solution of sulphurous acid.

$SO_4^{2-}(aq)$ $H_2SO_3(aq)$

$$\underset{A}{H_2SO_3(aq)} + \underset{B}{SO_4^{2-}(aq)} \xrightleftharpoons{>50\%} \underset{CB}{HSO_3^{1-}(aq)} + \underset{CA}{HSO_4^{1-}(aq)}$$

7. Ammonium fluoride is dissolved in water.

$\underset{SA}{NH_4^{1+}(aq)}$ $\underset{SB}{F^{1-}(aq)}$

$$\underset{A}{NH_4^{1+}(aq)} + \underset{B}{F^{1-}(aq)} \xrightleftharpoons{<50\%} \underset{CB}{NH_3(aq)} + \underset{CA}{HF(aq)}$$

8. Benzoic acid can be used as a starting material for the synthesis of aspirin. Write an equation for the reaction of benzoic acid with water.

$$\underset{A}{C_6H_5COOH(aq)} + \underset{B}{H_2O(l)} \xrightleftharpoons{<50\%} \underset{CB}{C_6H_5COO^{1-}(aq)} + \underset{CA}{H_3O^+(aq)}$$

9. In solution, the poisonous gas $H_2S(g)$, which is sometimes found in natural gas, reacts with carbonate ions.

$$\underset{A}{H_2S\,(aq)} + \underset{B}{CO_3^{2-}\,(aq)} \xrightleftharpoons{>50\%} \underset{CB}{HS^{1-}\,(aq)} + \underset{CA}{HCO_3^{1-}\,(aq)}$$

10. Rain water in an area near a tar sands plant could dissolve $SO_2(g)$. Write the equation that represents the reaction of this acid solution with sodium sulphate.

$H_2O(l) + SO_2(g) \rightarrow H_2SO_3(aq)$

$$\underset{A}{H_2SO_3(aq)} + \underset{B}{SO_4^{2-}\,(aq)} \xrightleftharpoons{>50\%} \underset{CB}{HSO_3^{1-}\,(aq)} + \underset{CA}{HSO_4^{1-}}$$

11. Vinegar, a dilute ethanoic acid solution, is used to neutralized some spilled lye (sodium hydroxide)

$$\underset{A}{CH_3COOH\,(aq)} + \underset{B}{OH^{1-}\,(aq)} \xrightleftharpoons{>50\%} \underset{CB}{CH_3COO^{1-}\,(aq)} + \underset{CA}{HOH\,(l)}$$

12. Sodium hydrogen carbonate may be used directly or in *gripe water* to neutralize excess stomach acid, $HCl_{(aq)}$.

$$\underset{B}{HCO_3^{1-}(aq)} + \underset{A}{H_3O^{+}(aq)} \xrightleftharpoons{>50\%} \underset{CA}{H_2CO_3(aq)} + \underset{CB}{H_2O(l)}$$

13. Sodium hydrogen carbonate is mixed with a solution of potassium dihydrogen phosphate.

$$\underset{B}{HCO_3^{1-}\,(aq)} + \underset{A}{H_2PO_4^{1-}(aq)} \xrightleftharpoons{<50\%} \underset{CB}{HPO_4^{2-}(aq)} + \underset{CA}{H_2CO_3(aq)}$$

Polyprotic Acids and Bases

For each of the following problems, write all reaction steps and the net reaction. Assume all reactions are quantitative.

1. Potassium hydroxide is continuously added to a carbonic acid solution.

$$H_2CO_3(aq) + OH^{1-}(aq) \longrightarrow \cancel{HCO_3^{1-}(aq)} + H_2O(l)$$
$$\cancel{HCO_3^{1-}(aq)} + OH^{1-}(aq) \longrightarrow CO_3^{2-}(aq) + H_2O(l)$$
$$H_2CO_3(aq) + 2OH^{1-}(aq) \longrightarrow CO_3^{2-}(aq) + 2H_2O(l)$$

2. Lithium hydroxide is continuously added to a citric acid solution, $C_3H_4OH(COOH)_3(aq)$. Citric acid can be abbreviated $H_3Ct(aq)$.

$$H_3Ct(aq) + OH^{1-}(aq) \longrightarrow \cancel{H_2Ct^{1-}(aq)} + H_2O(l)$$
$$\cancel{H_2Ct^{1-}(aq)} + OH^{1-}(aq) \longrightarrow \cancel{HCt^{2-}(aq)} + H_2O(l)$$
$$\cancel{HCt^{2-}(aq)} + OH^{1-}(aq) \longrightarrow Ct^{3-}(aq) + H_2O(l)$$
$$H_3Ct(aq) + 3OH^{1-}(aq) \longrightarrow Ct^{3-}(aq) + 3H_2O(l)$$

3. Potassium hydroxide is continuously added to sulphuric acid until no further reaction occurs.

$$H_2SO_4(aq) + OH^{1-}(aq) \longrightarrow \cancel{HSO_4^{1-}(aq)} + H_2O(l)$$
$$\cancel{HSO_4^{1-}}(aq) + OH^{1-}(aq) \longrightarrow SO_4^{2-}(aq) + H_2O(l)$$
$$H_2SO_4(aq) + 2OH^{1-}(aq) \longrightarrow SO_4^{2-}(aq) + 2H_2O(l)$$

4. A sodium carbonate solution is titrated with hydrobromic acid.

$$CO_3^{2-}(aq) + H_3O^+(aq) \longrightarrow \cancel{HCO_3^{1-}(aq)} + H_2O(l)$$
$$\cancel{HCO_3^{1-}}(aq) + H_3O^+(aq) \longrightarrow H_2CO_3(aq) + H_2O(l)$$
$$CO_3^{2-}(aq) + 2H_3O^+(aq) \longrightarrow H_2CO_3(aq) + 2H_2O(l)$$

5. A solution of sodium hydrogen phosphate is titrated with hydroiodic acid.

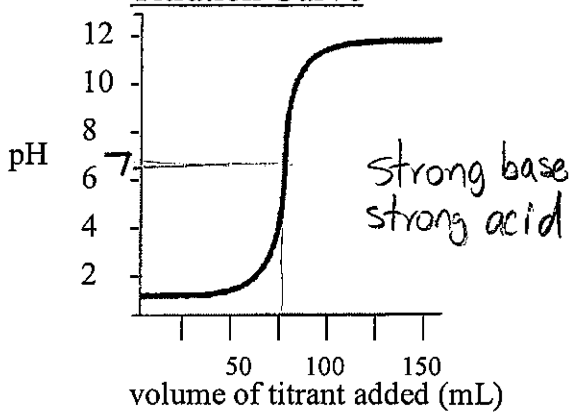
$$HPO_4^{2-}(aq) + H_3O^+(aq) \longrightarrow \cancel{H_2PO_4^{1-}}(aq) + H_2O(l)$$
$$\cancel{H_2PO_4^{1-}}(aq) + H_3O^+(aq) \longrightarrow H_3PO_4(aq) + H_2O(l)$$
$$HPO_4^{2-}(aq) + 2H_3O^+(aq) \longrightarrow H_3PO_4(aq) + 2H_2O(l)$$

pH Curves

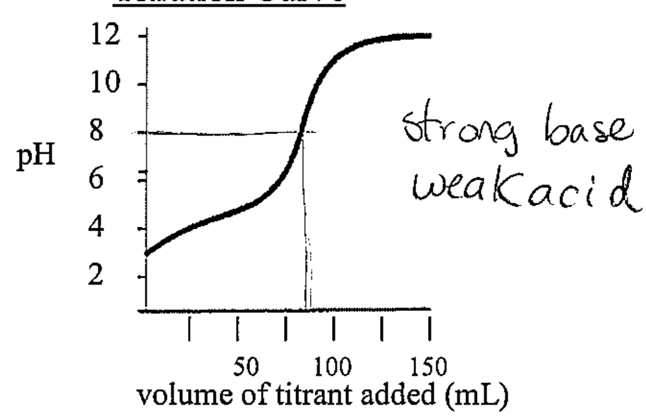
For each of the following curves state:

1. the titrant - substance that is added
2. what type of sample is being used eg) monoprotic acid, polybasic substance etc.
3. the endpoint pH value(s)
4. potential indicator for each endpoint
5. the equivalence point(s)

1. Titration Curve



2. Titration Curve



Curve 1:

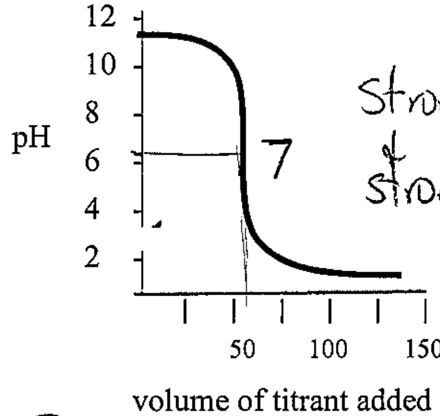
1. strong base
2. monoprotic strong acid
3. 7
4. bromothymol blue (best)
   phenol red
5. pH ~7 / 75mL added

Curve 2:

1. strong base
2. monoprotic weak acid
3. 8
4. phenol red
5. pH-8 / ~80mL

3.

Titration Curve

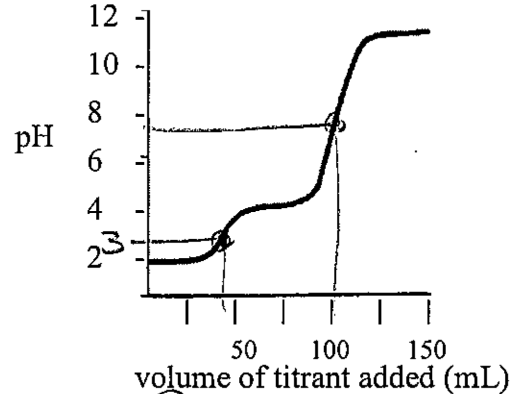


Strong acid + strong base

① strong acid
② monoprotic strong base
③ 7 ④ bromothymol blue
⑤ pH-7 ~55mL added

4.

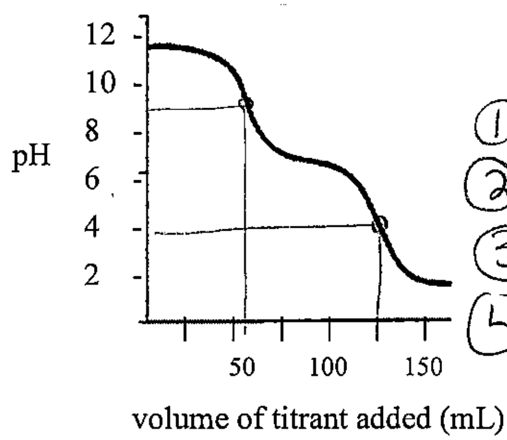
Titration Curve



① base
② diprotic acid ③ 3 and 7.6
④ 3 - orange IV / methyl orange
7.6 phenol red
⑤ ~3 + 7.6
~45mL added
about 100mL added

5.

Titration Curve



① acid
② diprotic base
③ @ ~9 + 4
④ phenolphthalein — methyl orange
⑤ ~55mL added ~125mL added

Buffers

1. One of the most important buffers in our bodies is the carbonic acid – hydrogen carbonate ion system.
   a) Write the equation for the reaction that occurs when a strong acid is added to this buffer.

$$HCO_3^{1-}(aq) + H_3O^+(aq) \rightleftharpoons H_2CO_3(aq) + H_2O(l)$$

   b) Write the equation for the reaction that occurs when a strong base is added to this buffer.

$$HCO_3^{1-}(aq) + OH^{1-}(aq) \rightleftharpoons CO_3^{2-}(aq) + H_2O(l)$$

2. Another important buffer in our bodies is the dihydrogen phosphate – hydrogen phosphate ion system.

   a) Write the equation for the reaction of this buffer with ammonia.

   $$H_2PO_4^{1-}(aq) + NH_3(g) \rightleftharpoons HPO_4^{2-}(aq) + NH_4^{1+}(aq)$$

   b) Write the equation for the reaction of this buffer with carbonic acid.

   $$HPO_4^{2-}(aq) + H_2CO_3(aq) \rightleftharpoons H_2PO_4^{1-}(aq) + HCO_3^{1-}(a$$