

BE READY END OF TERM 3 REVISION



تم تحميل هذا الملف من موقع المناهج الإماراتية

موقع المناهج ← المناهج الإماراتية ← الصف الحادي عشر المتقدم ← كيمياء ← الفصل الثالث ← ملفات متنوعة ← الملف

تاريخ إضافة الملف على موقع المناهج: 18:57:31 2026-06-15

ملفات اكتب للمعلم اكتب للطالب | اختبارات الكترونية | اختبارات | حلول | عروض بوربوينت | أوراق عمل
منهج انجليزي | ملخصات وتقارير | مذكرات وبنوك | الامتحان النهائي | للمدرس

المزيد من مادة
كيمياء:

إعداد: عبدالرحمن الشناوي

التواصل الاجتماعي بحسب الصف الحادي عشر المتقدم



صفحة المناهج
الإماراتية على
فيسبوك

الرياضيات

اللغة الانجليزية

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التربية الاسلامية

المواد على تلغرام

المزيد من الملفات بحسب الصف الحادي عشر المتقدم والمادة كيمياء في الفصل الثالث

محلول غير 1 P Adv 11 Grade Chemistry Revision 3 Term of End

1

إجابات 1 P Adv 11 Grade Chemistry Revision 3 Term of End

2

End of Term 3 Revision

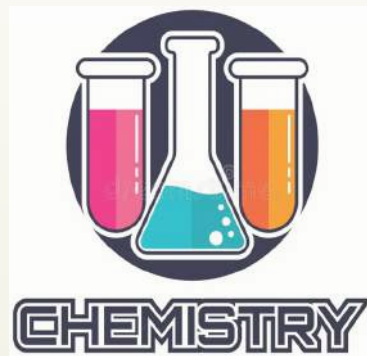
3

مراجعة هيكل الامتحان كراسة تدريبية

4

EOT3 - CHEMISTRY - ACIDS AND BASES - CHEMICAL EQUILIBRIUM

5



BE READY END OF TERM 3 REVISION CHEMISTRY GRADE 11 ADV P 2

COMPILED BY
MR.ABDULRAHMAN
ALSHENNAWI

مدير المجمع
أمجد الجبر

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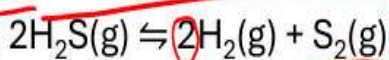
1	CHM.5.4.02.006.03 Calculate the equilibrium concentration given the value of the equilibrium constant and equilibrium concentration of reactants and products
2	CHM.5.4.02.006.04 Calculate the value of K_{eq} , K_c , or K_p when the concentration or pressure of reactants and products are given at a constant temperature
3	CHM.5.4.02.007.02 Write the solubility equilibrium expression for a given reaction equilibrium
4	CHM.5.4.02.008.01 Calculate the molar solubility (concentration) of a sparingly soluble ionic compound using the solubility product constant, K_{sp}
5	CHM.5.4.02.011.03 Explore the application of the solubility product constant in radiology while exploring its effects on human vital systems
6	CHM.5.4.02.007.03 Write the solubility for a given reaction equilibrium

IN-CLASS EXAMPLE

Use with Example Problem 4.

Problem

At a temperature of 1405 K, hydrogen sulfide, which has a foul odor resembling rotten eggs, decomposes to form hydrogen and diatomic sulfur. The equilibrium constant for the reaction is 2.27×10^{-3} .



KNOWN

- ✓ $K_{eq} = 2.27 \times 10^{-3}$
- ✓ $[\text{S}_2] = 0.0540 \text{ mol/L}$
- ✓ $[\text{H}_2\text{S}] = 0.184 \text{ mol/L}$

UNKNOWN

$[\text{H}_2] = ? \text{ mol/L}$

$$2.27 \times 10^{-3} = \frac{2^2 \times 0.054}{(0.184)^2} \quad (\text{shift solve})$$

What is the concentration of hydrogen gas $[\text{H}_2]$, if $[\text{S}_2] = 0.0540 \text{ mol/L}$ and $[\text{H}_2\text{S}] = 0.184 \text{ mol/L}$?

adsy

SOLVE FOR THE UNKNOWN

- State the equilibrium constant expression

$$K_{eq} = \frac{[H_2]^2 [S_2]}{[H_2S]^2} \Rightarrow \frac{[H_2]^2 [S_2]}{[S_2]} = \frac{K_{eq} [H_2S]^2}{[S_2]}$$

Solve the equation for $[H_2]$.

- Multiply both sides by $[H_2S]^2$. Divide both sides by $[S_2]$.

$$[H_2]^2 = K_{eq} \times \frac{[H_2S]^2}{[S_2]}$$

- Take the square root of both sides.

$$[H_2] = \sqrt{K_{eq} \times \frac{[H_2S]^2}{[S_2]}} \quad (H_2) = \sqrt{\frac{K_{eq} [H_2S]^2}{[S_2]}} = \sqrt{\frac{2.27 \times 10^{-3} \times (0.184)^2}{0.054}} = 0.0377 \text{ mol/L}$$

PRACTICE Problems

Module 1b Extra Practice
Lesson 3 part 1

18. At a certain temperature, $K_{eq} = 10.5$ for the equilibrium $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$.

Calculate the following concentrations:

- $[CO]$ in an equilibrium mixture containing 0.933 mol/L H_2 and 1.32 mol/L CH_3OH
- $[H_2]$ in an equilibrium mixture containing 1.09 mol/L CO and 0.325 mol/L CH_3OH
- $[CH_3OH]$ in an equilibrium mixture containing 0.0661 mol/L H_2 and 3.85 mol/L CO

$$K_{eq} = \frac{[CH_3OH]}{[CO] \times [H_2]^2} \Rightarrow [CH_3OH] = K_{eq} \times [CO] \times [H_2]^2$$

$$= 10.5 \times 3.85 \times 0.0661$$

$$= 0.177 \frac{\text{mol}}{\text{L}}$$

or $\frac{\text{M}}{\text{L}}$

18. At a certain temperature, $K_{eq} = 10.5$ for the equilibrium $\text{CO(g)} + 2\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_3\text{OH(g)}$.

Calculate the following concentrations.

- a. $[\text{CO}]$ in an equilibrium mixture containing 0.933 mol/L H_2 and 1.32 mol/L CH_3OH

$$K_{eq} = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

$$10.5 = \frac{(1.32)}{[\text{CO}](0.933)^2}$$

$$[\text{CO}] = 0.144\text{M}$$

- b. $[\text{H}_2]$ in an equilibrium mixture containing 1.09 mol/L CO and 0.325 mol/L CH_3OH

$$K_{eq} = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

$$10.5 = \frac{(0.325)}{(1.09)[\text{H}_2]^2}$$

$$[\text{H}_2] = 0.169\text{M}$$

- c. $[\text{CH}_3\text{OH}]$ in an equilibrium mixture containing 0.0661 mol/L H_2 and 3.85 mol/L CO

$$K_{eq} = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

$$10.5 = \frac{[\text{CH}_3\text{OH}]}{(3.85)(0.0661)^2}$$

$$[\text{CH}_3\text{OH}] = (10.5)(3.85)(0.0661)^2 = 0.177\text{M}$$

Answers

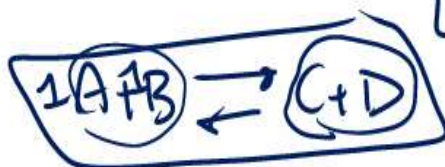
19. **Challenge** In a generic reaction $\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$, 1.00 mol of A and 1.00 mol of B are allowed to react in a 1-L flask until equilibrium is established. If the equilibrium concentration of A is 0.450 mol/L, what is the equilibrium concentration of each of the other substances? What is K_{eq} ?

$$[\text{A}]_{eq} = 0.450 \text{ mol/L} = [\text{B}]_{eq}$$

$$[\text{B}]_{eq} = ? = 0.45 \text{ mol/L}$$

$$[\text{C}]_{eq} = ?$$

$$[\text{D}] = ?$$



$$[\text{C}]_{eq} = [\text{D}]_{eq} = 0.55$$

$$K_{eq} = \frac{[\text{C}]_{eq}[\text{D}]_{eq}}{[\text{A}]_{eq}[\text{B}]_{eq}} = \frac{0.55 \times 0.55}{0.45 \times 0.45} = 1.49$$

$$[\text{A}]_i = 1 \frac{\text{mol}}{\text{L}}$$

$$[\text{A}]_{f(eq)} = 0.45$$

we used $1 - 0.45$

$$[\text{B}]_{eq} = 0.55 \text{ mol/L}$$

- 19. Challenge** In a generic reaction of $A + B \rightleftharpoons C + D$, 1.00 mol of A and 1.00 mol of B are allowed to react in a 1-L flask until equilibrium is established. If the equilibrium concentration of A is 0.450 mol/L, what is the equilibrium concentration of each of the other substances? What is K_{eq} ?

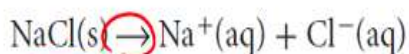
According to the stoichiometry of the equation, the concentration of B is 0.450M; C and D are $1.00 - 0.450 = 0.550M$.

$$K_{eq} = \frac{(0.550)(0.550)}{(0.450)(0.450)} = 1.49$$

Answers

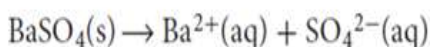
The Solubility Product Constant

Some ionic compounds, such as sodium chloride, dissolve readily in water, and some, such as barium sulfate ($BaSO_4$) barely dissolve at all. On dissolving, all ionic compounds dissociate into ions.

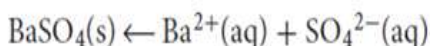


Sometimes low solubility is also important. Although barium ions are toxic to humans, patients must ingest barium sulfate prior to having an X ray of the digestive tract taken. Can patients safely ingest $BaSO_4$?

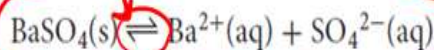
Barium sulfate dissociates in water according to this equation.



As soon as the first product ions form, the reverse reaction begins.

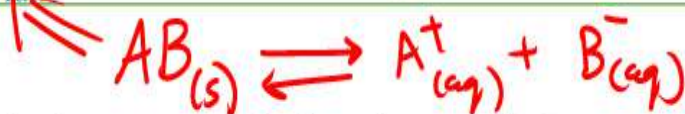


In time, equilibrium is established.





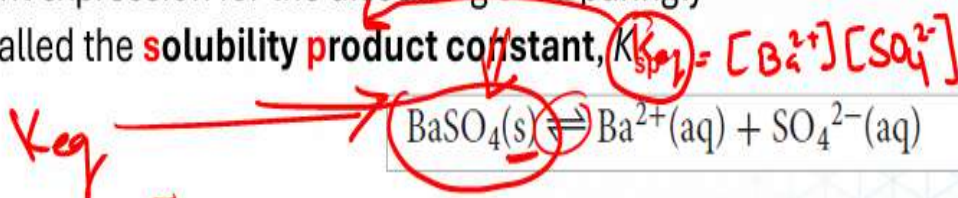
The Solubility Product Constant: P.101



Sparingly = low solubility

- Some ionic compounds dissolve readily in water, and some barely dissolve at all.

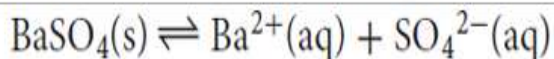
- The equilibrium constant expression for the dissolving of a sparingly soluble compound is called the **solubility product constant**, K_{sp}



- The **solubility product constant expression** " K_{sp} " is the product of the concentrations of the dissolved ions, each raised to the power equal to the coefficient of the ion in the chemical equation.

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- Experimentally determined values can be used to determine the solubility of a sparingly soluble compound.



Now you can write the solubility product constant expression for the dissolving of barium sulfate (BaSO_4) in water. The K_{sp} for the process is 1.1×10^{-10} at 298 K.

$$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

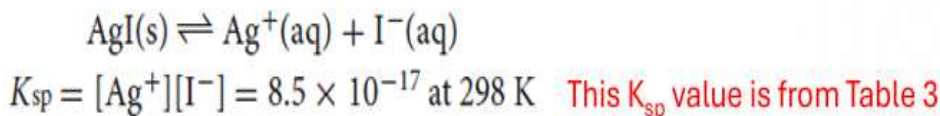
- K_{sp} can be used to predict whether a precipitate will form when any two ionic solutions are mixed.

Table 17.3

Solubility Product Constants at 298 K

Compound	K_{sp}	Compound	K_{sp}	Compound	K_{sp}
Carbonates		Halides		Hydroxides	
BaCO_3	2.6×10^{-9}	CaF_2	3.5×10^{-11}	$\text{Al}(\text{OH})_3$	4.6×10^{-33}
CaCO_3	3.4×10^{-9}	PbBr_2	6.6×10^{-6}	$\text{Ca}(\text{OH})_2$	5.0×10^{-6}
CuCO_3	2.5×10^{-10}	PbCl_2	1.7×10^{-5}	$\text{Cu}(\text{OH})_2$	2.2×10^{-20}
PbCO_3	7.4×10^{-14}	PbF_2	3.3×10^{-8}	$\text{Fe}(\text{OH})_3$	4.9×10^{-17}
MgCO_3	6.8×10^{-6}	PbI_2	9.8×10^{-9}	$\text{Fe}(\text{OH})_3$	2.8×10^{-39}
Ag_2CO_3	8.5×10^{-12}	AgCl	1.8×10^{-10}	$\text{Mg}(\text{OH})_2$	5.6×10^{-12}
ZnCO_3	1.5×10^{-10}	AgBr	5.4×10^{-13}	$\text{Zn}(\text{OH})_2$	3×10^{-17}
Hg_2CO_3	3.6×10^{-17}	AgI	8.5×10^{-17}	Sulfates	
Chromates		Phosphates		BaSO_4	1.1×10^{-10}
BaCrO_4	1.2×10^{-10}	AlPO_4	9.8×10^{-21}	CaSO_4	4.9×10^{-5}
PbCrO_4	2.3×10^{-13}	$\text{Ca}_3(\text{PO}_4)_2$	2.1×10^{-33}	PbSO_4	2.5×10^{-8}
Ag_2CrO_4	1.1×10^{-12}	$\text{Mg}_3(\text{PO}_4)_2$	1.0×10^{-24}	Ag_2SO_4	1.2×10^{-5}

Suppose you wish to determine the solubility of silver iodide (AgI) in mol/L at 298 K. The equilibrium equation and solubility product constant expression are as follows.



It is convenient to let s represent the solubility of AgI, that is, the number of moles of AgI that dissolves in one liter of solution. The equation indicates that for every mole of AgI that dissolves, an equal number of moles of Ag^+ ions forms in solution. Therefore, $[\text{Ag}^+]$ equals s . Every Ag^+ has an accompanying I^- ion, so $[\text{I}^-]$ also equals s . Substituting s for $[\text{Ag}^+]$ and $[\text{I}^-]$, the K_{sp} expression becomes the following.

$$[\text{Ag}^+][\text{I}^-] = (s)(s) = s^2 = 8.5 \times 10^{-17}$$

$$s = \sqrt{8.5 \times 10^{-17}} = 9.2 \times 10^{-9} \text{ mol/L}$$

The solubility of AgI is $9.2 \times 10^{-9} \text{ mol/L}$ at 298 K.

chemistry4all



Figure 17 Greater definition is possible in a gastrointestinal X-ray when patients drink a thick mixture containing barium sulfate. Barium sulfate is a poisonous substance, but it has such low solubility that only a minimal amount can dissolve in the patient's body.

Now you can write the solubility product constant expression for the dissolving of barium sulfate (BaSO_4) in water. The K_{sp} for the process is 1.1×10^{-10} at 298 K.

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

$$s = [\text{Ba}^{2+}] = [\text{SO}_4^{2-}]$$

$$K_{sp} = s \cdot s = s^2 \Rightarrow s = \sqrt{K_{sp}}$$

$$s = \sqrt{1.1 \times 10^{-10}} = 1.048 \times 10^{-5} \text{ M}$$

SOLVE FOR THE UNKNOWN

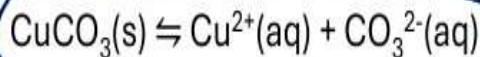
Use with Example Problem 5.

Page 104

Problem

Use the K_{sp} value from Table 3 to calculate the solubility in mol/L of copper(II) carbonate (CuCO_3) at 298 K.

- State the balanced chemical equation for the solubility equilibrium.



- State the solubility product constant expression.

$$K_{sp} = [\text{Cu}^{2+}][\text{CO}_3^{2-}] = 2.5 \times 10^{-10}$$

- Relate $[\text{Cu}^{2+}]$ and $[\text{CO}_3^{2-}]$ to the solubility of CuCO_3 , s .

$$s = [\text{Cu}^{2+}] = [\text{CO}_3^{2-}]$$

- Substitute s for $[\text{Cu}^{2+}]$ and $[\text{CO}_3^{2-}]$ in the expression for K_{sp} .

$$(s)(s) = s^2 = 2.5 \times 10^{-10}$$

SOLVE FOR THE UNKNOWN (continued)

- Solve for s , and calculate the answer.

$$s = \sqrt{2.5 \times 10^{-10}} = 1.6 \times 10^{-5} \text{ mol/L}$$

The molar solubility of CuCO_3 in water at 298 K is 1.6×10^{-5} mol/L.

KNOWN

UNKNOWN

$$K_{sp}(\text{CuCO}_3) = 2.5 \times 10^{-10}$$

$$s = ? \text{ mol/L}$$

IN-CLASS EXAMPLE

Use with Example Problem 5.

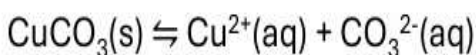
Page 104

Problem

Use the K_{sp} value from Table 3 to calculate the solubility in mol/L of copper(II) carbonate (CuCO_3) at 298 K.

SOLVE FOR THE UNKNOWN

- State the balanced chemical equation for the solubility equilibrium.



- State the solubility product constant expression.

$$K_{sp} = [\text{Cu}^{2+}][\text{CO}_3^{2-}] = 2.5 \times 10^{-10}$$

- Relate $[\text{Cu}^{2+}]$ and $[\text{CO}_3^{2-}]$ to the solubility of CuCO_3 , s .

$$s = [\text{Cu}^{2+}] = [\text{CO}_3^{2-}]$$

- Substitute s for $[\text{Cu}^{2+}]$ and $[\text{CO}_3^{2-}]$ in the expression for K_{sp} .

$$(s)(s) = s^2 = 2.5 \times 10^{-10}$$

SOLVE FOR THE UNKNOWN (continued)

- Solve for s , and calculate the answer.

$$s = \sqrt{2.5 \times 10^{-10}} = 1.6 \times 10^{-5} \text{ mol/L}$$

The molar solubility of CuCO_3 in water at 298 K is 1.6×10^{-5} mol/L.

KNOWN

UNKNOWN

$$K_{sp}(\text{CuCO}_3) = 2.5 \times 10^{-10}$$

$$s = ? \text{ mol/L}$$

CALCULATING MOLAR SOLUBILITY

IN-CLASS EXAMPLE

SOLVE FOR THE UNKNOWN
 (continued)

- Solve for s , and calculate the answer.

$$s = \sqrt{2.5 \times 10^{-10}} = 1.6 \times 10^{-5} \text{ mol/L}$$

The molar solubility of CuCO_3 in water at 298 K is $1.6 \times 10^{-5} \text{ mol/L}$.

EVALUATE THE ANSWER

The K_{sp} value has two significant figures, so the answer is correctly expressed with two digits. As predicted, the molar solubility of CuCO_3 is approximately 10^{-5} mol/L .

PRACTICE Problems

Extra Practice Page 988 and glencoe.com

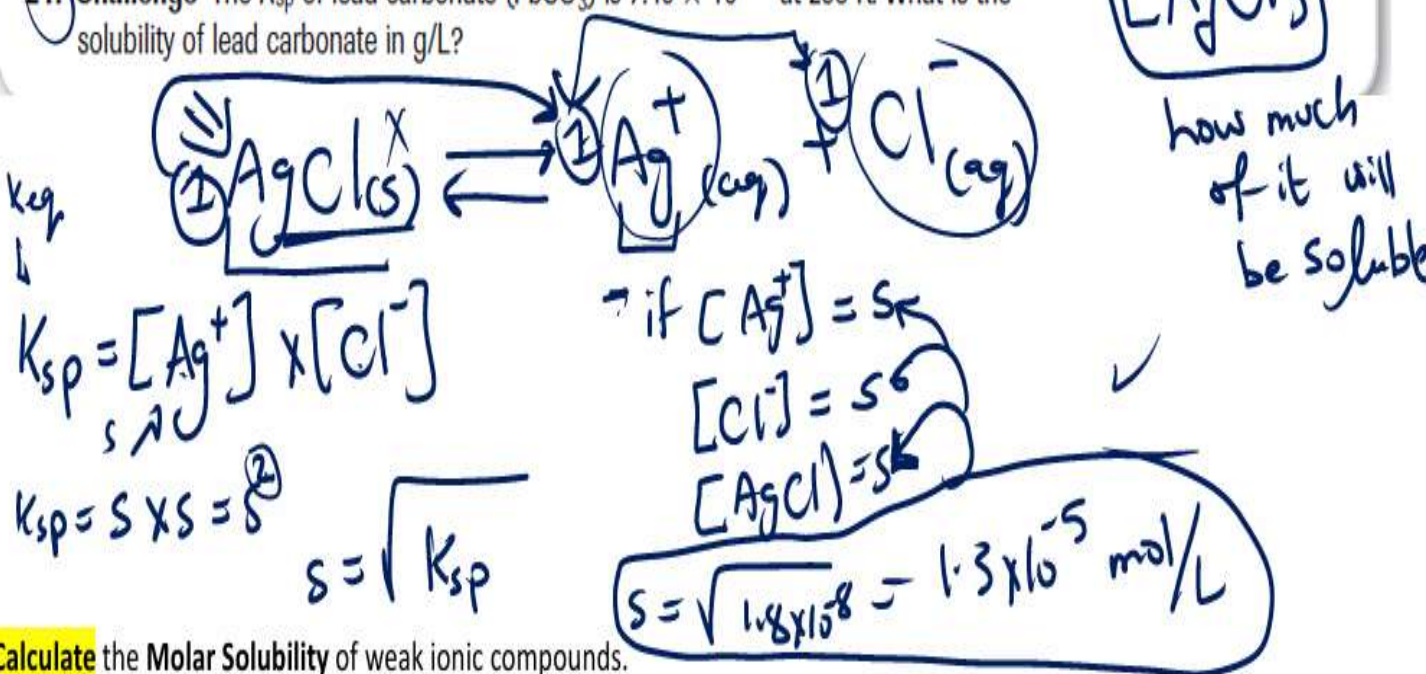
20. Use the data in Table 17.3 to calculate the solubility in mol/L of the following ionic compounds at 298 K.

a. PbCrO_4

b. AgCl

c. CaCO_3

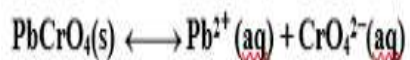
21. Challenge The K_{sp} of lead carbonate (PbCO_3) is 7.40×10^{-14} at 298 K. What is the solubility of lead carbonate in g/L?



Calculate the Molar Solubility of weak ionic compounds.

20. Use the data in Table 3 to calculate the solubility in mol/L of the following ionic compounds at 298 K.

a. PbCrO_4



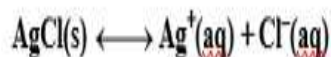
s mol/L dissolves s mol/L s mol/L

$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{CrO}_4^{2-}]$$

$$2.33 \times 10^{-13} = (s)(s) = s^2$$

$$s = \sqrt{2.33 \times 10^{-13}} = 4.8 \times 10^{-7} \text{ M}$$

b. AgCl



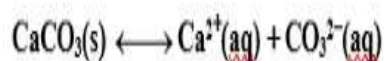
s mol/L dissolves s mol/L s mol/L

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = (s)(s) = s^2$$

$$s = \sqrt{1.8 \times 10^{-10}} = 1.3 \times 10^{-5} \text{ M}$$

c. CaCO_3



s mol/L dissolves s mol/L s mol/L

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

$$3.4 \times 10^{-9} = (s)(s) = s^2$$

$$s = \sqrt{3.4 \times 10^{-9}} = 5.8 \times 10^{-5} \text{ M}$$

21. **Challenge** The K_{sp} of lead(II) carbonate (PbCO_3) is 7.40×10^{-14} at 298 K. What is the solubility of lead (II) carbonate in g/L?

$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{CO}_3^{2-}] = 7.40 \times 10^{-14}$$

$$(s)(s) = 7.40 \times 10^{-14}$$

$$s = \sqrt{7.40 \times 10^{-14}} = 2.72 \times 10^{-7} \text{ M}$$

$$s = 2.72 \times 10^{-7} \text{ mol/L} \times 267.2 \text{ g/mol}$$

$$= 7.27 \times 10^{-5} \text{ g/L}$$

Answers

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CALCULATING ION CONCENTRATION

IN-CLASS EXAMPLE

Use with Example Problem 6.

Problem Page 105

Magnesium hydroxide is a white solid obtained from seawater and used in the formulation of many medications, in particular those whose function is to neutralize excess stomach acid.

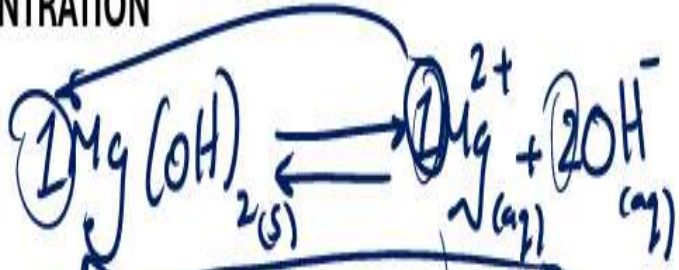
Determine the hydroxide ion concentration in a saturated solution of $Mg(OH)_2$ at 298 K. The K_{sp} equals 5.6×10^{-12} .

KNOWN

$$K_{sp} = 5.6 \times 10^{-12}$$

UNKNOWN

$$[OH^-] = ? \text{ mol/L}$$



$$K_{sp} = [Mg^{2+}] \times [OH^-]^2$$

let's say $[Mg^{2+}] = x$

$$[OH^-] = 2x$$

solubility $[Mg(OH)_2] = x$

if you find x, you find all of them

$$K_{sp} = x \cdot (2x)^2 = x \cdot 4x^2 = \frac{4x^3}{4}$$

$$x^3 = \frac{K_{sp}}{4} \Rightarrow x = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{5.6 \times 10^{-12}}{4}} = \boxed{}$$

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SOLVE FOR THE UNKNOWN (continued)

- Use your calculator to determine the cube root.

$$x = [\text{Mg}^{2+}] = \sqrt[3]{1.4 \times 10^{-12}} = 1.1 \times 10^{-4} \text{ mol/L}$$

Multiply $[\text{Mg}^{2+}]$ by 2 to obtain $[\text{OH}^-]$.

$$[\text{OH}^-] = 2[\text{Mg}^{2+}] = 2(1.1 \times 10^{-4} \text{ mol/L}) = 2.2 \times 10^{-4} \text{ mol/L}$$

$$[\text{Mg}^{2+}] = x = 1.1 \times 10^{-4}$$

$$[\text{Mg}(\text{OH})_2] = x = 1.1 \times 10^{-4}$$

IN-CLASS EXAMPLE

Use with Example Problem 6.

Problem Page 105

Magnesium hydroxide is a white solid obtained from seawater and used in the formulation of many medications, in particular those whose function is to neutralize excess stomach acid. Determine the hydroxide ion concentration in a saturated solution of $\text{Mg}(\text{OH})_2$ at 298 K. The K_{sp} equals 5.6×10^{-12} .

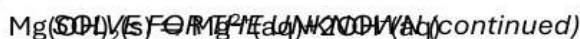
KNOWN

$$K_{\text{sp}} = 5.6 \times 10^{-12}$$

UNKNOWN

$$[\text{OH}^-] = ? \text{ mol/L}$$

- State the equation for the solubility equilibrium.



- State the K_{sp} expression.

$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 5.6 \times 10^{-12}$$

Let $x = [\text{Mg}^{2+}]$. Because there are two OH^- ions for every Mg^{2+} ion, $2x = [\text{OH}^-]$.

- Substitute $x = [\text{Mg}^{2+}]$ and $2x = [\text{OH}^-]$

$$(x)(2x)^2 = 5.6 \times 10^{-12}$$

- Square the terms.

$$(x)(4)(x)^2 = 5.6 \times 10^{-12}$$

- Combine the terms.

$$4x^3 = 5.6 \times 10^{-12}$$

- Divide.

$$x^3 = \frac{5.6 \times 10^{-12}}{4} = 1.4 \times 10^{-12}$$

SOLVE FOR THE UNKNOWN (continued)

- Use your calculator to determine the cube root.

$$x = [\text{Mg}^{2+}] = \sqrt[3]{1.4 \times 10^{-12}} = 1.1 \times 10^{-4} \text{ mol/L}$$

Multiply $[\text{Mg}^{2+}]$ by 2 to obtain $[\text{OH}^-]$.

$$[\text{OH}^-] = 2[\text{Mg}^{2+}] = 2(1.1 \times 10^{-4} \text{ mol/L}) = 2.2 \times 10^{-4} \text{ mol/L}$$

EVALUATE THE ANSWER

The given K_{sp} has two significant figures, so the answer is correctly stated with two digits. As predicted, $[\text{OH}^-]$ is about 10^{-4} mol/L .

PRACTICE Problems

Module 16 lesson

Extra Practice Page 988 and glencoe.com

22. Use K_{sp} values from Table 17.3 to calculate the following.

3 part 2

a. $[\text{Ag}^+]$ in a solution of AgBr at equilibrium

b. $[\text{F}^-]$ in a saturated solution of CaF_2

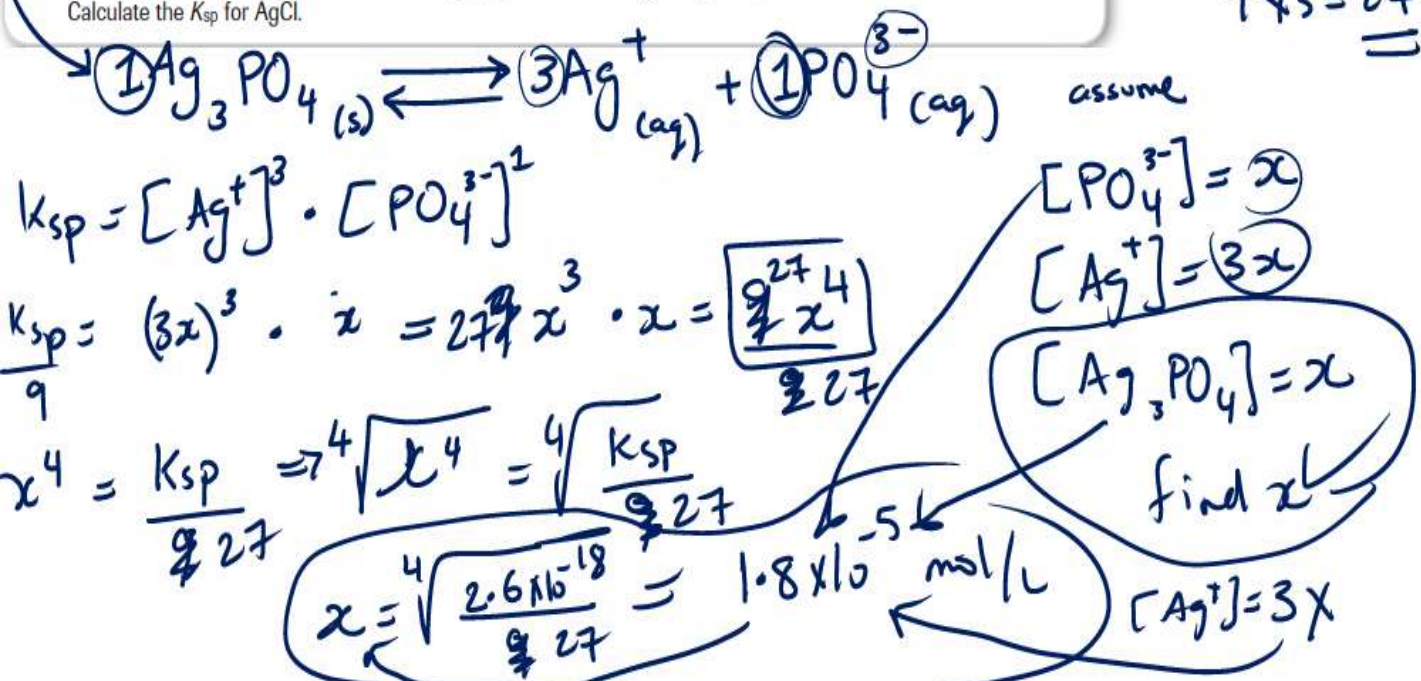
c. $[\text{Ag}^+]$ in a solution of Ag_2CrO_4 at equilibrium

$$[\text{Ag}_3\text{PO}_4] = ?!$$

23. Calculate the solubility of Ag_3PO_4 ($K_{sp} = 2.6 \times 10^{-18}$).

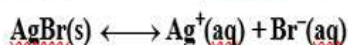
24. Challenge The solubility of silver chloride (AgCl) is $1.86 \times 10^{-4} \text{ g/100 g}$ of H_2O at 298 K. Calculate the K_{sp} for AgCl .

$$3^3 = 3 \times 3 \times 3 = 9 \times 3 = 27$$



22. Use the K_{sp} values from Table 3 to calculate the following.

a. $[Ag^+]$ in a solution of $AgBr$ at equilibrium



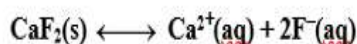
s mol/L dissolves s mol/L s mol/L

$$K_{sp} = [Ag^+][Br^-]$$

$$5.4 \times 10^{-13} = (s)(s) = s^2$$

$$s = \sqrt{5.4 \times 10^{-13}} = 7.3 \times 10^{-7} M = [Ag^+]$$

b. $[F^-]$ in a saturated solution of CaF_2



s mol/L dissolves s mol/L $2s$ mol/L

$$s = \frac{1}{2}[F^-]$$

$$K_{sp} = [Ca^{2+}][F^-]^2$$

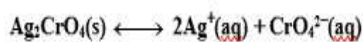
$$3.5 \times 10^{-11} = (s)(2s)^2 = 4s^3$$

$$s = \sqrt[3]{\frac{3.5 \times 10^{-11}}{4}} = 2.1 \times 10^{-4} M$$

$$\frac{1}{2}[F^-] = 2.1 \times 10^{-4} M$$

$$[F^-] = 4.2 \times 10^{-4} M$$

c. $[Ag^+]$ in a solution of Ag_2CrO_4 at equilibrium



s mol/L dissolves $2s$ mol/L s mol/L

$$s = \frac{1}{2}[Ag^+]$$

$$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$$

$$1.1 \times 10^{-12} = (2s)^2(s) = 4s^3$$

$$s = \sqrt[3]{\frac{1.1 \times 10^{-12}}{4}} = 6.5 \times 10^{-5} M$$

$$\frac{1}{2}[Ag^+] = 6.5 \times 10^{-5} M$$

$$[Ag^+] = 1.3 \times 10^{-4} M$$

Answers

23. Calculate the solubility of Ag_3PO_4 ($K_{sp} = 2.6 \times 10^{-18}$)

$$K_{sp} = [Ag^+]^3[PO_4^{3-}] = 2.6 \times 10^{-18}$$

$$[PO_4^{3-}] = s, [Ag^+] = 3s$$

$$(3s)^3(s) = (27s^3)(s) = 27s^4 = 2.6 \times 10^{-18}$$

$$s = \sqrt[4]{\frac{2.6 \times 10^{-18}}{27}} = 1.8 \times 10^{-5} \text{ mol/L}$$

24. Challenge The solubility of silver chloride ($AgCl$) is 1.86×10^{-4} g/100 g of H_2O at 298 K. Calculate the K_{sp} for $AgCl$.

$$s = \frac{1.86 \times 10^{-4} \text{ g}}{100 \text{ g } H_2O} \times \frac{1000 \text{ g } H_2O}{1 \text{ L}}$$

$$= 1.86 \times 10^{-3} \text{ g/L}$$

$$s = \frac{1.86 \times 10^{-3} \text{ g}}{1 \text{ L}} \times \frac{1 \text{ mol}}{143.4 \text{ g}} = 1.30 \times 10^{-5} \text{ mol/L}$$

$$K_{sp} = [Ag^+][Cl^-] = (s)(s)$$

$$K_{sp} = (1.30 \times 10^{-5} \text{ mol/L})(1.30 \times 10^{-5} \text{ mol/L})$$

$$= 1.7 \times 10^{-10}$$

Easy Chemistry 4all

19. Challenge In a generic reaction $A + B \rightleftharpoons C + D$, 1.00 mol of A and 1.00 mol of B are allowed to react in a 1-L flask until equilibrium is established. If the equilibrium concentration of A is 0.450 M , what is the equilibrium concentration of each of the other substances? What is K_{eq} ?

Method: ICE
initial change equilibrium

	A	B	C	D
I	1 M	1 M	0	0
C	-x	-x	+x	+x
E	$1-x$	$1-x$	x	x

$0.45 = 1 - x$
 $x = 0.55$

Eq:

$$[A] = 0.45 = [B]$$

$$[C] = [D] = 0.55$$

$$K_{eq} = \frac{[C][D]}{[A][B]}$$

$$= \frac{0.55^2}{0.45^2} = 1.5$$

19. Challenge In a generic reaction $A + B \rightleftharpoons C + D$, 1.00 mol of A and 1.00 mol of B are allowed to react in a 1-L flask until equilibrium is established. If the equilibrium concentration of A is 0.450 mol/L, what is the equilibrium concentration of each of the other substances? What is K_{eq} ?

ICE method

Reaction: $A + B \rightleftharpoons C + D$

Initial concentrations: $[A] = 1.0\text{ M}$, $[B] = 1.0\text{ M}$, $[C] = 0$, $[D] = 0$

Equilibrium concentrations: $[A]_{eq} = 0.45\text{ M}$, $[B]_{eq} = 0.45\text{ M}$, $[C]_{eq} = 0.55\text{ M}$, $[D]_{eq} = 0.55\text{ M}$

ICE Table:

Initial	Change	Final Eq
1	-x	1-x
1	-x	1-x
0	+x	x
0	+x	x

Equilibrium concentrations: $[A]_{eq} = [B]_{eq} = 1 - x = 0.45$

Solving for x: $x = 0.55$

Equilibrium constant: $K_{eq} = \frac{[C]_{eq}[D]_{eq}}{[A]_{eq}[B]_{eq}} = \frac{0.55^2}{0.45^2} = 1.49$

Easy Chemistry4all

21. Challenge The K_{sp} of lead carbonate ($PbCO_3$) is 7.40×10^{-14} at 298 K. What is the solubility of lead carbonate in g/L?

$M_m = 267.21\text{ g/mol}$

Dissociation: $PbCO_3(s) \rightleftharpoons Pb^{2+}(aq) + CO_3^{2-}(aq)$

$K_{sp} = [Pb^{2+}][CO_3^{2-}] = s \cdot s = s^2$

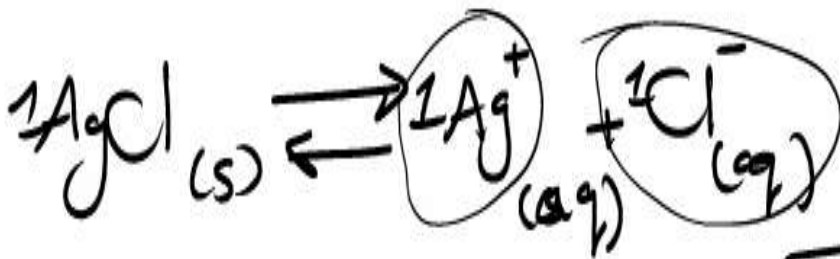
Solving for s: $s = \sqrt{7.40 \times 10^{-14}} = 2.72 \times 10^{-7}\text{ mol/L}$

Converting to g/L: $2.72 \times 10^{-7}\text{ mol/L} \times 267.21\text{ g/mol} = 7.268 \times 10^{-5}\text{ g/L}$

ICE Table:

	$PbCO_3(s)$	Pb^{2+}	CO_3^{2-}
I	y	0	0
C	-s	+s	+s
E	y-s	s	s

Calculate the Molar Solubility of weak ionic compounds.



$$K_{sp} = S^2$$

24. Challenge The solubility of silver chloride (AgCl) is 1.86×10^{-4} g/100 g of H_2O at 298 K. Calculate the K_{sp} for AgCl.

$$M_m = 143.32 \text{ g/mol}$$

given

$$\left(\frac{\text{mol}}{\text{L}} \right)$$

unknown K_{sp} :-

$$S = 1.86 \times 10^{-4} \text{ g AgCl}$$

Density

$$D_{\text{H}_2\text{O}} = \left(\frac{1 \text{ g}}{1 \text{ mL}} \right)$$

$$D = \frac{m}{V} \Rightarrow V = \frac{m}{D}$$

$$100 \text{ g H}_2\text{O}$$

$$100 \text{ g H}_2\text{O} \rightarrow 100 \text{ mL H}_2\text{O}$$

$$V = \frac{100 \text{ g}}{1 \text{ g/mL}} = 100 \text{ mL} \xrightarrow{\div 1000} 0.1 \text{ L}$$

$$1 \text{ g AgCl} \rightarrow \text{mol}$$

$$n = \frac{m}{M_m} = \frac{1.86 \times 10^{-4}}{143.32} = 1.3 \times 10^{-6} \text{ mol}$$

$$S = \frac{1.3 \times 10^{-6} \text{ mol}}{0.1 \text{ L}} = 1.3 \times 10^{-5} \frac{\text{mol}}{\text{L}}$$

$$K_{sp} = S^2 = (1.3 \times 10^{-5})^2 = 1.69 \times 10^{-10}$$

مجمع زايد التعليمي
الخوانيج



The reaction : $\text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{Cl}_2(\text{g})$

reaches equilibrium at 900 K.

$$K_{eq} = 8.20 \times 10^{-2}$$

If the equilibrium concentrations of CO and Cl_2 are 0.150 mol/L,

what is the equilibrium concentration of COCl_2 ?

$\text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{Cl}_2(\text{g})$: الفاعل :

يصل إلى حالة الاتزان عند درجة حرارة 900 K .

$$K_{eq} = 8.20 \times 10^{-2}$$

إذا كانت تراكيز الاتزان لـ CO و Cl_2 هي 0.150 mol/L ،

لما تركز الاتزان لـ COCl_2 ؟

$$K_{eq} = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]}$$

0.274 mol/L

$$8.2 \times 10^{-2} = \frac{0.15^2}{x}$$

0.183 mol/L

3.65 mol/L

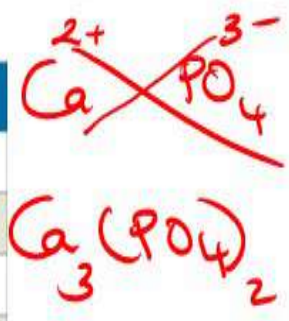
$$x = 0.274 \text{ M}$$

0.550 mol/L

مجمع زايد التعليمي
الخوانيج

Easy Chemistry 4th

Solubility Product Constants at 298 K					
Compound	K_{sp}	Compound	K_{sp}	Compound	K_{sp}
Carbonates		Halides		Hydroxides	
BaCO ₃	2.6×10^{-9}	CaF ₂	3.5×10^{-11}	Al(OH) ₃	4.6×10^{-33}
CaCO ₃	3.4×10^{-9}	PbBr ₂	6.6×10^{-6}	Ca(OH) ₂	5.0×10^{-6}
CuCO ₃	2.5×10^{-10}	PbCl ₂	1.7×10^{-5}	Cu(OH) ₂	2.2×10^{-20}
PbCO ₃	7.4×10^{-14}	PbF ₂	3.3×10^{-8}	Fe(OH) ₂	4.9×10^{-17}
MgCO ₃	6.8×10^{-6}	PbI ₂	9.8×10^{-9}	Fe(OH) ₃	2.8×10^{-39}
Ag ₂ CO ₃	8.5×10^{-12}	AgCl	1.8×10^{-10}	Mg(OH) ₂	5.6×10^{-12}
ZnCO ₃	1.5×10^{-10}	AgBr	5.4×10^{-13}	Zn(OH) ₂	3×10^{-17}
Hg ₂ CO ₃	3.6×10^{-17}	AgI	8.5×10^{-17}	Sulfates	
Chromates		Phosphates		BaSO ₄	1.1×10^{-10}
BaCrO ₄	1.2×10^{-10}	AlPO ₄	9.8×10^{-21}	CaSO ₄	4.9×10^{-5}
PbCrO ₄	2.3×10^{-13}	Ca ₃ (PO ₄) ₂	2.1×10^{-33}	PbSO ₄	2.5×10^{-8}
Ag ₂ CrO ₄	1.1×10^{-12}	Mg ₃ (PO ₄) ₂	1.0×10^{-24}	Ag ₂ SO ₄	1.2×10^{-5}

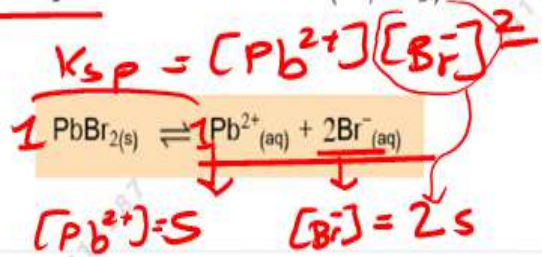


What is the solubility in mol/L of lead bromide **PbBr₂** at 298 K?

ما ذائبة بروميد الرصاص PbBr₂ عند 298 K (بوحدة mol/L)؟

if $K_{sp} = 6.6 \times 10^{-6}$

إذا كان $K_{sp} = 6.6 \times 10^{-6}$



~~$S/2S^2$~~

0.024

0.018

0.030

0.012

$K_{sp} = S \cdot (2S)^2 = S \cdot 4S^2 = 4S^3$

$6.6 \times 10^{-6} = 4S^3$

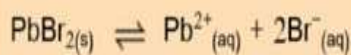
$S = 0.0118 \approx 0.012$

What is the solubility in mol/L of lead bromide PbBr_2 at 298 K?

ما ذائبية بروميد الرصاص PbBr_2 عند 298 K (بوحدة mol/L)؟

if $K_{sp} = 6.6 \times 10^{-6}$

إذا كان $K_{sp} = 6.6 \times 10^{-6}$



0.024

0.018

0.030

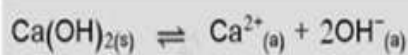
0.012

If the K_{sp} of calcium hydroxide Ca(OH)_2 is 5.32×10^{-6} at 298 K. What is the solubility of calcium hydroxide in mol/L?

إذا كانت قيمة K_{sp} لهيدروكسيد الكالسيوم Ca(OH)_2 تساوي

5.32×10^{-6} عند 298 K . فما ذائبية هيدروكسيد الكالسيوم بوحدة

mol/L ؟



3.30×10^{-8}

1.60×10^{-3}

2.80×10^{-7}

1.10×10^{-2}

$$K_{sp} = 4s^3$$

$$5.32 \times 10^{-6} = 4s^3$$

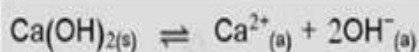
$$s = 0.01099$$

$\approx$

$$1.099 \times 10^{-2} \approx 1.10 \times 10^{-2}$$

If the K_{sp} of calcium hydroxide Ca(OH)_2 is 5.32×10^{-6} at 298 K. What is the solubility of calcium hydroxide in mol/L?

إذا كانت قيمة K_{sp} لهيدروكسيد الكالسيوم Ca(OH)_2 تساوي 5.32×10^{-6} عند 298 K . فما ذائبية هيدروكسيد الكالسيوم بوحدة mol/L?



$$3.30 \times 10^{-8}$$

$$1.60 \times 10^{-3}$$

$$2.80 \times 10^{-7}$$

$$1.10 \times 10^{-2}$$

If the K_{sp} of magnesium hydroxide Mg(OH)_2 is 5.60×10^{-12} at 298 K. What is the solubility of magnesium hydroxide in mol/L?

إذا كانت قيمة K_{sp} لهيدروكسيد المغنيسيوم Mg(OH)_2 تساوي 5.60×10^{-12} عند 298 K . فما ذائبية هيدروكسيد المغنيسيوم بوحدة mol/L?



$$K_{sp} = 4s^3$$

$$5.6 \times 10^{-12} = 4s^3$$

$$s = 1.1186 \times 10^{-4}$$

$$3.45 \times 10^{-5}$$

$$1.12 \times 10^{-4}$$

$$4.73 \times 10^{-6}$$

$$2.52 \times 10^{-4}$$

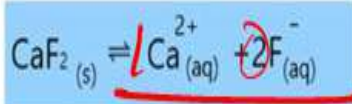
What is the value of K_{sp} for CaF_2 compound if the

ما قيمة K_{sp} للمركب CaF_2 إذا كانت ذائبية فلوريد الكالسيوم CaF_2 تساوي

solubility of calcium fluoride CaF_2 is $2.06 \times 10^{-4} \text{ M}$

$2.06 \times 10^{-4} \text{ M}$ عند درجة حرارة 298 K ؟

at 298 K?



$$K_{sp} = 4s^3$$

$$4.2 \times 10^{-4}$$

$$= 4 (2.06 \times 10^{-4})^3$$

$$1.8 \times 10^{-7}$$

$$8.8 \times 10^{-12}$$

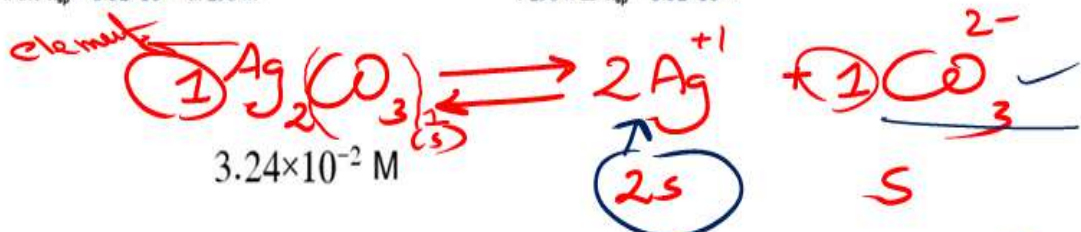
$$3.5 \times 10^{-11}$$

$$3.496 \times 10^{-11}$$

easy chemistry

What is the concentration of silver ion $[\text{Ag}^+]$ in Ag_2CO_3 solution at equilibrium if the $K_{sp} = 8.52 \times 10^{-12}$ at 298 K?

ما تركيز أيون الفضة $[\text{Ag}^+]$ في محلول Ag_2CO_3 عند الاتزان إذا كان $K_{sp} = 8.52 \times 10^{-12}$ عند 298 K ؟



$$3.24 \times 10^{-2} \text{ M}$$

$$8.81 \times 10^{-4} \text{ M}$$

$$2.57 \times 10^{-4} \text{ M}$$

$$1.29 \times 10^{-4} \text{ M}$$

$$K_{sp} = 4s^3 \Rightarrow 8.52 \times 10^{-12} = 4s^3$$

$$s = 1.285 \times 10^{-4} \text{ M}$$

$$[\text{Ag}^+] = 2s$$

$$= 2 \times 1.285 \times 10^{-4} = 2.57 \times 10^{-4}$$

7	CHM.5.3.04.001.01 List six general properties of aqueous acids (taste, color of indicators, reaction with metals, metal carbonates and bases, and electrical conductivity)
8	CHM.5.3.04.001.03 Differentiate among acidic, basic and neutral solutions (in terms of the relative amounts of hydrogen ions and hydroxide ions)
9	CHM.5.3.04.001.09 Describe the amphoteric behavior of water, H_2O , and ammonia, NH_3 (Using chemical equation, particulate diagram and space-filling models)
10	CHM.5.3.04.001.10 Distinguish among monoprotic, diprotic and triprotic acids using ionization equations, examples and particulate diagrams
11	CHM.5.3.04.001.08 Define acids and bases according to Brønsted-Lowry theory, indicating the acid, base, conjugate acid, conjugate base and conjugate acid-base pairs, when chemical equations, formula or space-filling models are given
12	CHM.5.3.04.001.11 Define acids and bases according to Lewis theory

What are the differences between **acids** and **bases**!!??

Properties of Acids and Bases

Acids taste sour (حامض). Bases taste bitter (مر) and feel slippery. ✓

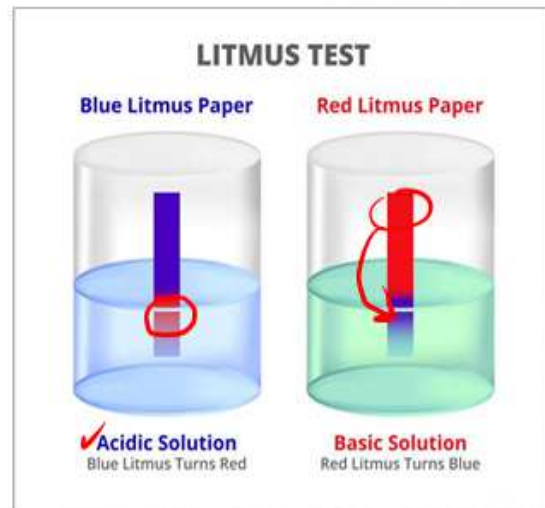
Litmus paper



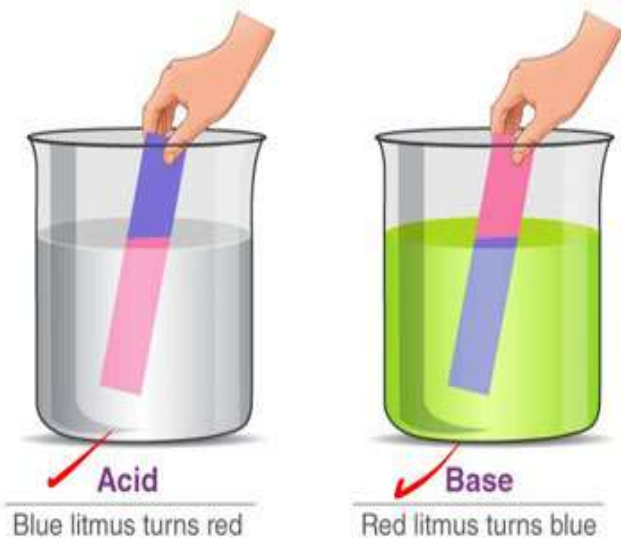
Acids and bases are **conductors of electricity**.

Acids turn **blue** litmus **red**.

Bases turn **red** litmus **blue**.



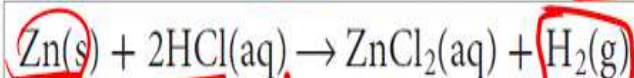
More on Litmus if you
interested!



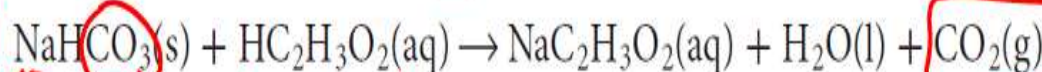
<https://en.wikipedia.org/wiki/Litmus>

- Acids and bases can be identified by their reactions with some metals and metal carbonates. CO_3^{2-}

- Magnesium and zinc react with acids to produce hydrogen gas (H_2).



- Metal carbonates and hydrogen carbonates also react with aqueous solutions of acids to produce carbon dioxide (CO_2) gas.



Limestone:
a rock composed primarily of CaCO_3 .

- Geologists identify limestone (CaCO_3) because it produces bubbles of carbon dioxide when exposed to hydrochloric acid.

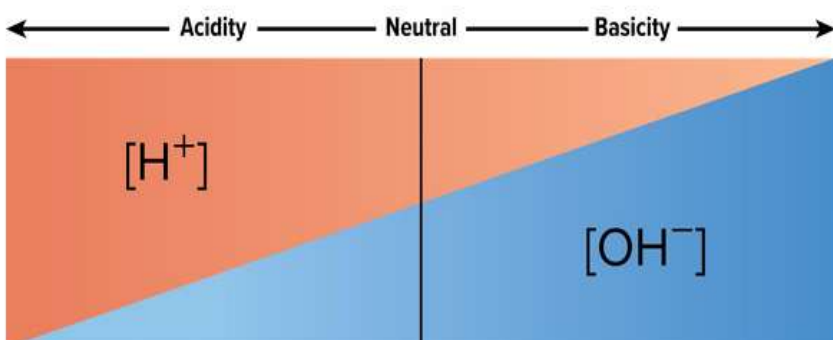


Properties of Acids and Bases

- All water solutions contain hydrogen ions (H^+) and hydroxide ions (OH^-).
- An **acidic solution** contains **more hydrogen ions** than hydroxide ions.
- A **basic solution** contains **more hydroxide ions** than hydrogen ions.

Acids and bases are **conductors of electricity** because they produce ions (H^+) and (OH^-) when they dissolve in water.

"Ions in water = conducting electricity"

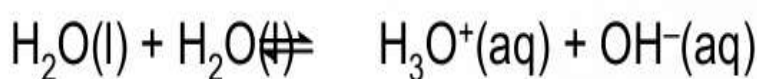
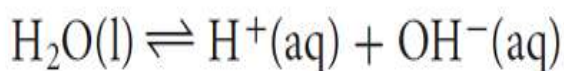


Remember: $[\text{H}^+]$ means the concentration (molarity) of H^+ in M or mol/L

Easy Chemistry: all
The usual solvent for acids and bases is **water**.

- Water produces **equal numbers of hydrogen (H^+) and hydroxide (OH^-) ions** in a process called **self-ionization (التأين الذاتي)**.

Ionization= produces ions



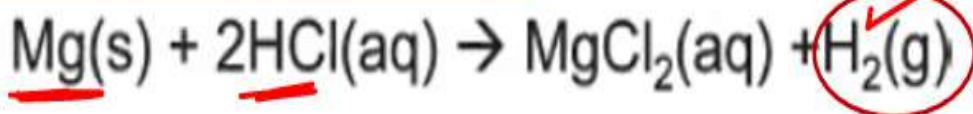
- The hydronium ion, H_3O^+** , is a hydrogen ion that has a water molecule attached to it by a covalent bond. **The symbols H^+ and H_3O^+ are used interchangeably.**

1. Which of these is **NOT** a property of a **base**?

- ☐ A tastes bitter
- ☐ B conducts electricity
- ☐ C turns red litmus blue
- ☒ D reacts with zinc to form H_2 gas

Based on the below reaction, HCl is considered an acid.
Why?

Hint: Mg is Magnesium (Metal)

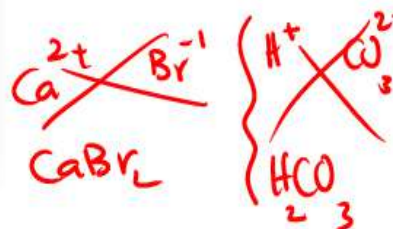


- Because Magnesium and zinc react with **acids** to produce hydrogen gas (H_2).

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

- Write balanced equations for the reactions between the following.
 - aluminum and sulfuric acid
 - calcium carbonate and hydrobromic acid
- Challenge** Write the net ionic equation for the reaction in Question 1b.

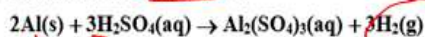


Practice Problems

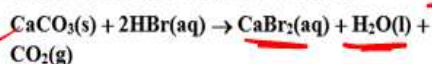
page 559

- Write balanced equations for reactions between the following.

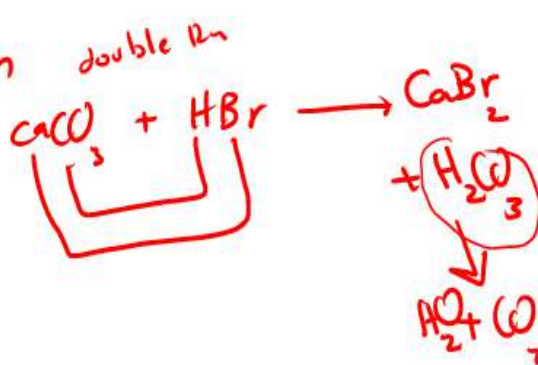
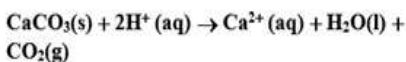
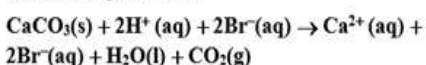
a. ^{Metal} aluminum and ^{H₂SO₄} sulfuric acid



b. calcium carbonate and hydrobromic acid

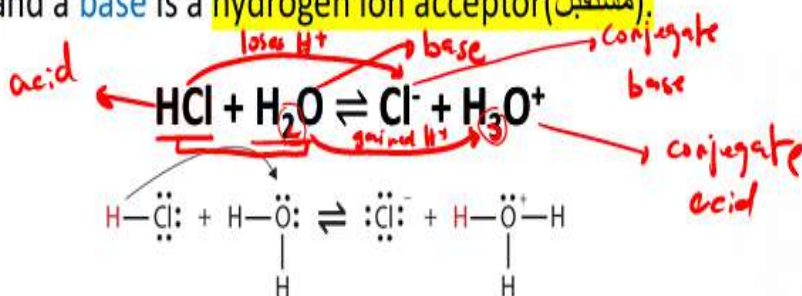


- Challenge** Write the net ionic equation for the reaction in question 1b.



The Brønsted-Lowry Model

- The Brønsted-Lowry model of acids and bases states that an **acid** is a hydrogen ion donor (فائق), and a **base** is a hydrogen ion acceptor (مستقبل).



- The Brønsted-Lowry model is a **more inclusive model** than the Arrhenius model.

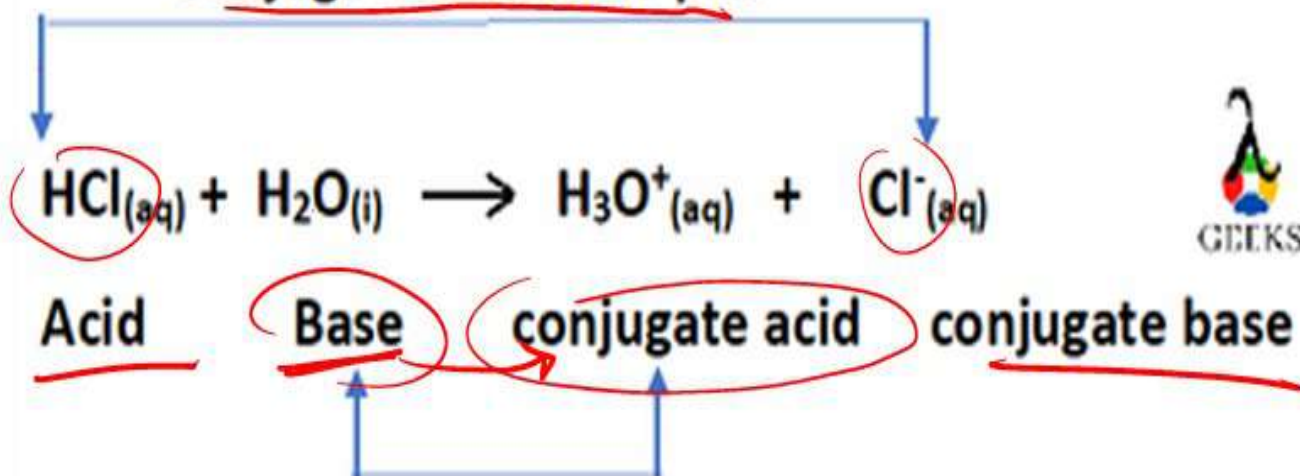
Conjugate=المرافق

- A **conjugate acid** is the species produced when a **base** accepts a **hydrogen ion**.
- A **conjugate base** is the species produced when an **acid** donates a **hydrogen ion**.
- A **conjugate acid-base pair** consists of two substances related to each other by donating and accepting a single hydrogen ion.

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الخوانيج

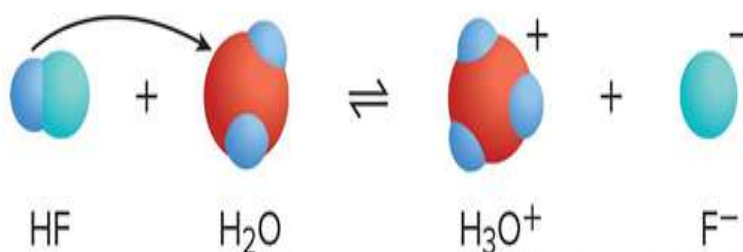
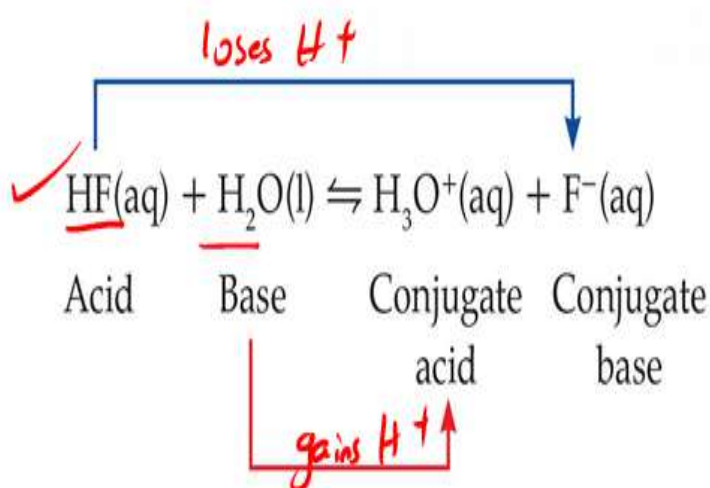
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Conjugate acid-base pair



Conjugate acid-base pair

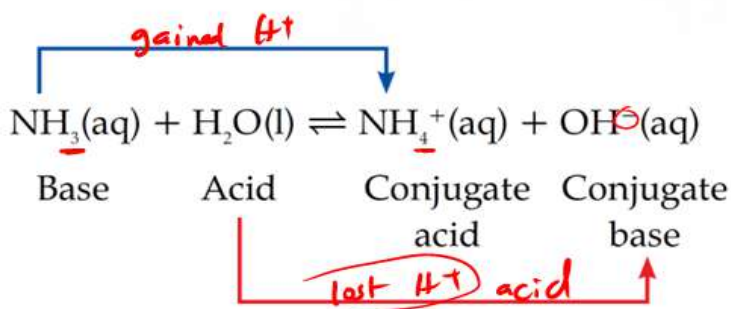
Hydrogen fluoride donates a hydrogen ion to a water molecule, so hydrogen fluoride is an acid.



4. What are **bases** defined as in the Brønsted-Lowry model?

- A OH^- producers
- B H^+ acceptors**
- C electron-pair acceptors
- D electron-pair donors

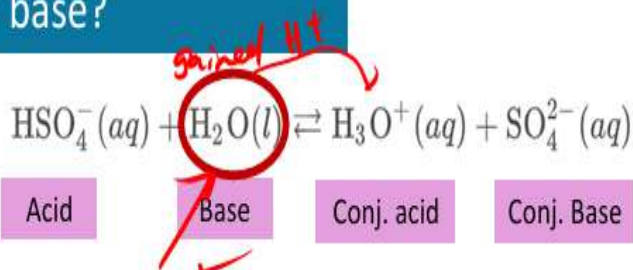
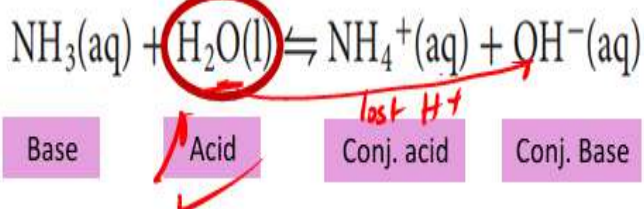
- Ammonia accepts a hydrogen ion, so it is a base.



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الخوانيج



Compare water H_2O in both reactions, is it acid or base?



So, water can act as **both an acid and base** depends on what other substance are in the solution.

❖ Water and other substances that can **act as both acids and bases** are said to be amphoteric.

An amphoteric substance is one that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.

2. Amphoteric substances such as water _____.

☒ always act as acids

☒ always act as bases

☒ never act as acids or bases

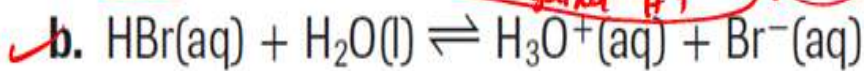
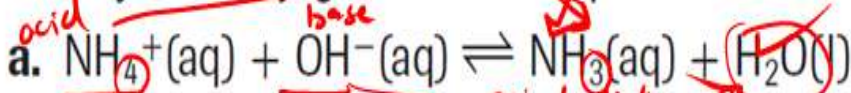
☐ D can act as either acids or bases

CORRECT

PRACTICE Problems

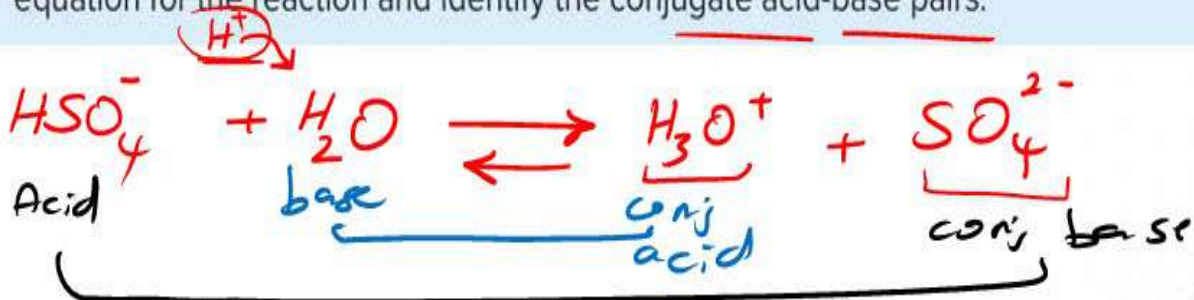
Extra Practice Page 989 and glencoe.com

3. Identify the conjugate acid-base pairs in each reaction.



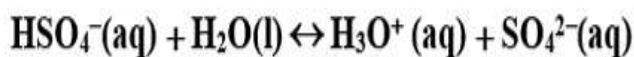
Acid	Conjugate base	Base	Conjugate acid
a. NH_4^+	NH_3	OH^-	H_2O
b. HBr	Br^-	H_2O	H_3O^+
c. H_2O	OH^-	CO_3^{2-}	HCO_3^-

4. **CHALLENGE** The products of an acid-base reaction are H_3O^+ and SO_4^{2-} . Write a balanced equation for the reaction and identify the conjugate acid-base pairs.



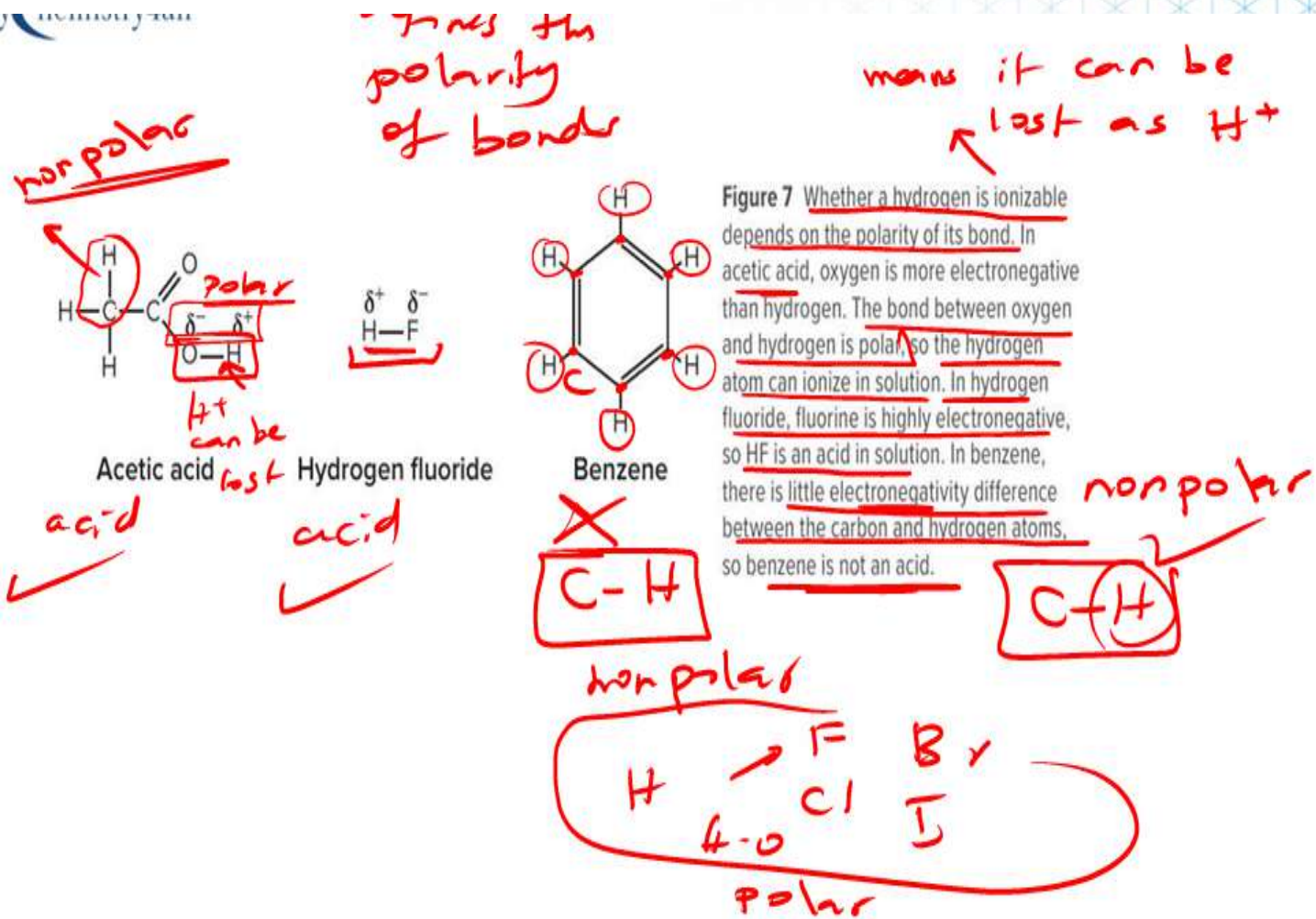
4. CHALLENGE The products of an acid-base reaction are H_3O^+ and SO_4^{2-} . Write a balanced equation for the reaction and identify the conjugate acid-base pairs.

4. Challenge The products of an acid-base reaction are H_3O^+ and SO_4^{2-} . Write a balanced equation for the reaction and identify the conjugate acid-base pairs.

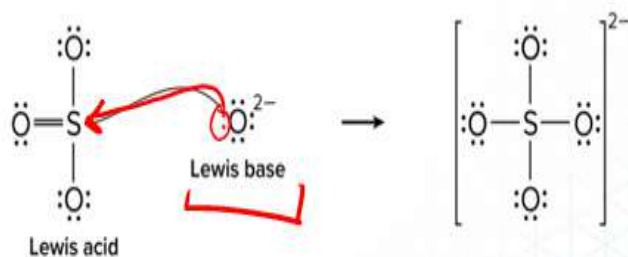
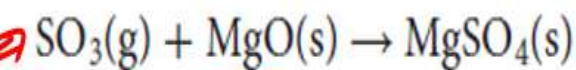
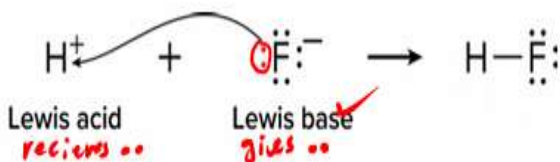


Reactant base: H_2O ; conjugate acid: H_3O^+

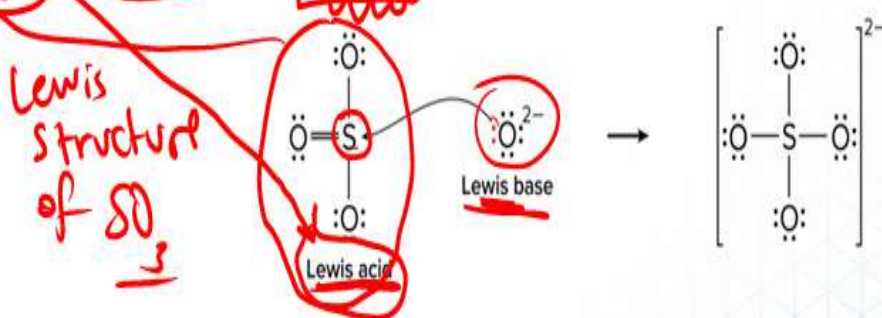
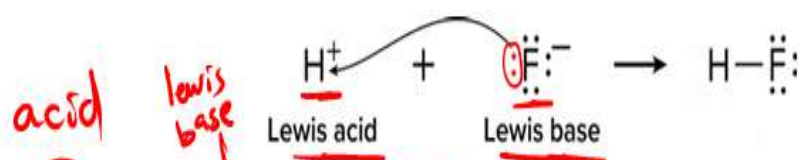
Reactant acid: HSO_4^- ; conjugate base: SO_4^{2-}



- According to the **Lewis model**, a **Lewis acid** is an **electron-pair acceptor**, and a **Lewis base** is an **electron-pair donor**. *accepts $\cdot\cdot e^-$ / loses gives $\cdot\cdot e^-$*
- The Lewis model includes all the substances classified as Brønsted-Lowry acids and bases and many more.



- According to the **Lewis model**, a **Lewis acid** is an **electron-pair acceptor**, and a **Lewis base** is an **electron-pair donor**.
- The Lewis model includes all the substances classified as Brønsted-Lowry acids and bases and many more.



you might they will draw it for you

Three Models for Acids and Bases

Model	Acid Definition	Base Definition
Arrhenius	H^+ producer	OH^- producer
✓ Brønsted-Lowry	H^+ donor	H^+ acceptor
✓ Lewis	electron-pair acceptor	electron-pair donor

stry4all

For each description below, write *acid* if it tells about a property of an acid or *base* if it tells about a property of a base. If the property does not apply to either an acid or a base, write *neither*. If it applies to both an acid and a base, write *both*.

- Both 1. Can turn litmus paper a different color
Red to blue Base
- Acid 2. Reacts with certain metals
Blue to red Acid
 $Acid + metal \rightarrow salt$
- Acid 3. Contains more hydrogen ions than hydroxide ions
 $[H^+] > [OH^-]$
 H^+ OH^- H_2O
- Base 4. Feels slippery
- Acid 5. Reacts with carbonates
 CO_3^{2-}
 $Acid + Carbonates \rightarrow CO_2 + H_2O$
- Neither 6. Feels rough
- Neither 7. Contains equal numbers of hydrogen and hydroxide ions
 $[H^+] = [OH^-]$
 Neutral (water)
- Base 8. Tastes bitter
- Acid 9. Tastes sour

What gas is produced during the reaction between sodium carbonate and acetic acid solution

ما الغاز الناتج خلال تفاعل كربونات الصوديوم الهيدروجينية مع المحلول المائي لحمض الأسيتيك؟

$CO_2 (g)$

Which of the following is a property of basic solutions?

أي مما يأتي من خصائص محاليل القواعد؟

- | | | |
|---|--|--|
| a | React with metals like zinc to produce hydrogen gas | تتفاعل مع الفلزات مثل الزنك لإنتاج غاز الهيدروجين |
| b | React with metal carbonates to produce carbon dioxide (CO_2) gas | تتفاعل مع كربونات الفلزات لإنتاج غاز ثاني أكسيد الكربون (CO_2) |
| c | Taste sour | طعمها لاذع |
| d | Taste bitter and feel slippery | مرارة المذاق وزلقة اللمس |
- Base* ✓

Which of the following chemical equations represents the reaction that geologists use to identify limestone rocks from other rocks?

أي المعادلات الكيميائية الآتية تمثل التفاعل الذي يستخدمه الجيولوجيون للتعرف على الصخور الجيرية من بقية الصخور؟

$CaCO_3$

- | | | |
|---|--|---|
| a | $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$ | ✓ |
| b | $NaHCO_{3(s)} + HC_2H_3O_{2(aq)} \rightarrow NaC_2H_3O_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$ | |
| c | $Cu_{(s)} + 4HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2H_2O_{(l)} + 2NO_{2(g)}$ | |
| d | $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$ | |

Which of the following is a property of acids?

أي مما يأتي يُعتبر من خصائص الأحماض؟

Turn red litmus paper blue

B

تُحول ورقة تباع الشمس الحمراء إلى اللون الأزرق

☐

Feel slippery

B

زلفة الملمس

☐

React with zinc to produce hydrogen gas

A

تتفاعل مع الخارصين لتنتج غاز الهيدروجين

☐

It tastes bitter

B

طعمها مر

☐

Which of the following chemical equations represents a reaction

between the aqueous solution of an acid and metal hydrogen

أي المعادلات الكيميائية التالية تمثل تفاعل بين المحلول المائي

لحمض وكربونات الفلز الهيدروجينية؟

metal

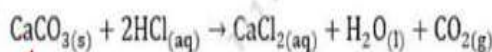
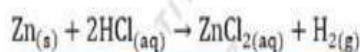
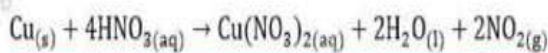
bicarbonate $\Rightarrow$ metal

H carbonate

carbonate?

CO_3

HCO_3



metal carbonate

Which of the following is correct?

$[OH^-] > [H^+]$ B
 أي مما يأتي صحيح؟
 $[H^+] = [OH^-]$ Neutral

- Acidic**
- ☒ a In basic solution $[H^+] > [OH^-]$ في المحلول القاعدي يكون $[H^+] > [OH^-]$ ☐
- acidic**
- ☒ b In neutral solution $[H^+] > [OH^-]$ في المحلول المتعادل يكون $[H^+] > [OH^-]$ ☐
- Basic**
- ☒ c In acidic solution $[H^+] > [OH^-]$ في المحلول الحمضي يكون $[H^+] > [OH^-]$ ☐
- Basic**
- ☒ d In acidic solution $[H^+] < [OH^-]$ في المحلول الحمضي يكون $[H^+] < [OH^-]$ ☐

Definitions

Table 2 Three Models for Acids and Bases

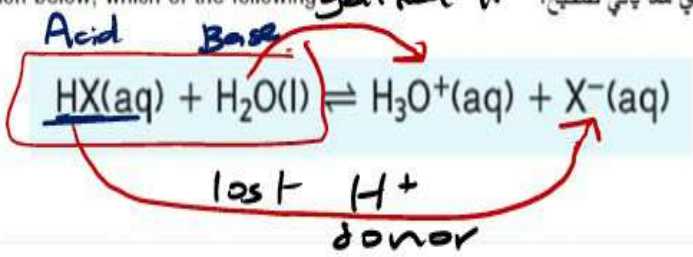
Model	Acid Definition	Base Definition
Arrhenius	H^+ producer in water	OH^- producer in water
Brønsted-Lowry B-L	H^+ donor gives H^+	H^+ acceptor accepts H^+
<u>Lewis</u>	electron-pair acceptor gains (e^-)	electron-pair donor gives e^-

Easy chemistry4all

Using (H^+)

B-L

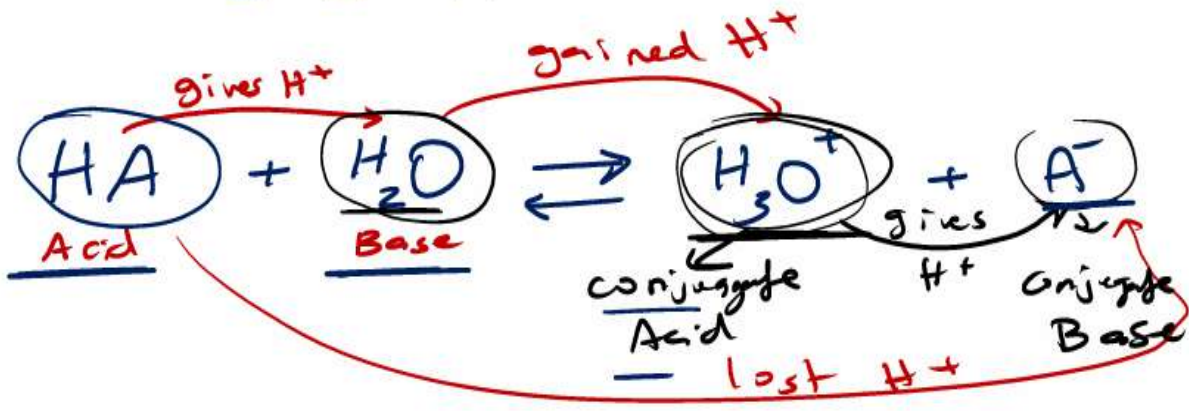
In the reaction equation below, which of the following is true?
 في معادلة التفاعل أدناه، أي مما يأتي صحيح؟



- ☒ a HX donates hydrogen ion to water H_2O يمنح HX أيون هيدروجين للماء H_2O 0
- ☐ b H_2O is a Bronsted-Lowry acid يُعتبر H_2O من أحماض برونشتد - لوري 0
- ☐ c HX is a Bronsted-Lowry base يُعتبر HX من قواعد برونشتد - لوري 0
- ☐ d HX accepts a hydrogen ion from water H_2O يستقبل HX أيون هيدروجين من الماء H_2O 0

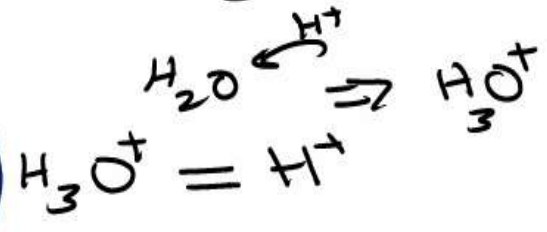
YouTube
 chemistry4all

B-L Acids & Bases



HA/A^- conj A-B pair
 H_2O/H_3O^+ conj A-B pair

H_3O^+ is called "Hydronium ion"



YouTube

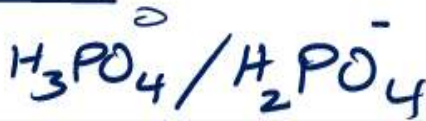
4all



Which of the following is considered a conjugate

أي مما يأتي يُعتبر زوج حمض قاعدة مرافق؟

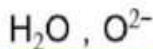
acid -base pair?



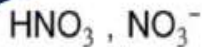
(a)



(b)



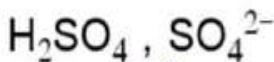
(c)



Acid

conj - Base

(d)



nemisuy4an

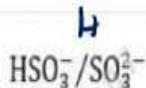
Which of the following is **not** a conjugate acid- base pair?

أي مما يلي **ليس** زوج حمض قاعدة مرافق؟

(a)



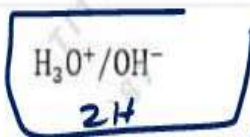
(b)



(c)



(d)



2H



Amphoteric

look for
polyprotic

Diprotic
triprotic

acids

Which of the following is amphoteric substance?

HPO_4^{2-} H^+ H_3PO_4 PO_4^{2-}

Amphoteric

H_3PO_4 H_2PO_4^- HPO_4^{2-} PO_4^{3-}

أي مما يلي يعتبر مادة أمفوتيرية

Amphoteric - a substance that can be an acid or a base in different reactions.

eg $\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$

H_2O is amphoteric

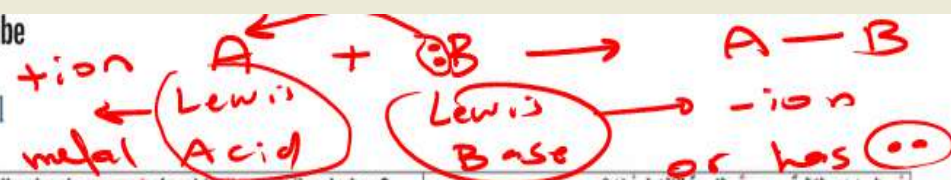
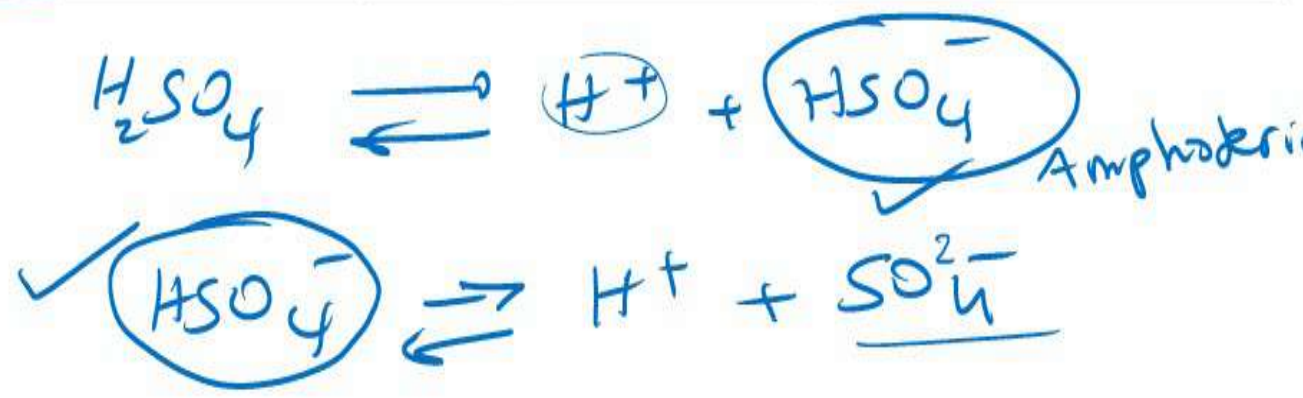
$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{NH}_4^+$

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 الخوانيج

Amphoteric

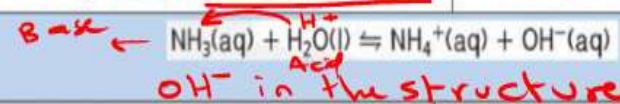
Table 2 Three Models for Acids and Bases

Model	Acid Definition	Base Definition
✓ Arrhenius	H ⁺ producer	OH ⁻ producer
Brønsted-Lowry ✓	H ⁺ donor	H ⁺ acceptor
<u>Lewis</u>	<u>electron-pair acceptor</u> .. accepts	<u>electron-pair donor</u> .. loses



Which of the following is correct about the reaction below?

أي العبارات التالية صحيحة بالنسبة للتفاعل أدناه؟



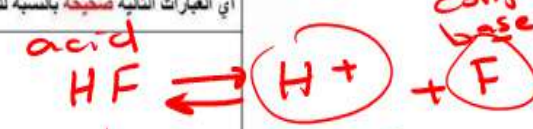
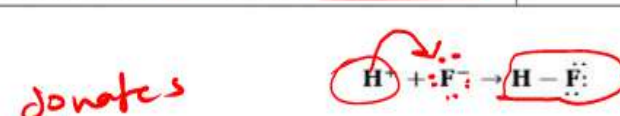
KOH

- ☒ A. NH₃ is considered as Arrhenius base
- ☒ B. H₂O is considered as a Bronsted - Lowry acid
- ☒ C. NH₃ accepts an electron pair from H₂O
- ☒ D. H₂O accepts H⁺ ion in the forward reaction

- A. يُعتبر NH₃ قاعدة أرهينوس
- B. يُعتبر H₂O حمض برونستد - لوري
- C. NH₃ يستقبل زوج الإلكترونات من H₂O
- D. H₂O يستقبل أيون H⁺ في التفاعل الأمامي

Which of the following is correct about the reaction below?

أي العبارات التالية صحيحة بالنسبة للتفاعل أدناه؟

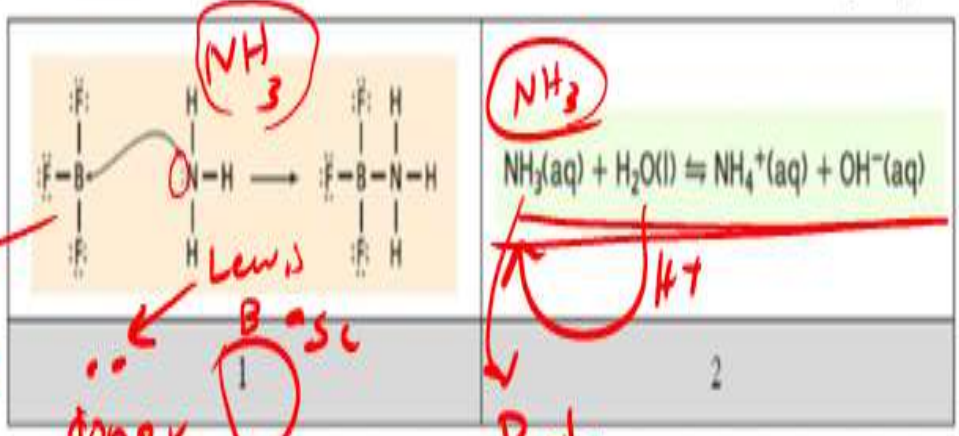


- ☒ A. F⁻ ion accepts an electron pair
- ☒ B. F⁻ ion is considered as acceptor of hydrogen ion
- ☒ C. H⁺ ion is considered as Arrhenius base
- ☒ D. H⁺ ion donates an electron pair to F⁻ ion

- A. يستقبل أيون F⁻ زوج إلكترونات
- B. يُعتبر F⁻ مستقبل لأيون الهيدروجين
- C. يُعتبر أيون H⁺ قاعدة أرهينوس
- D. يمنح أيون H⁺ زوج من الإلكترونات إلى أيون F⁻

Which of the following is true?

أي مما يأتي صحيح؟



X

The ammonia NH_3 in reaction 2 is an electron pair acceptor

تعتبر الأمونيا NH_3 في التفاعل 2 مستقبل زوج الإلكترونات

✓

The ammonia NH_3 in reaction 1 is a Lewis base

تعتبر الأمونيا NH_3 في التفاعل 1 قاعدة لويس

X

The ammonia NH_3 in reaction 2 is a Bronsted-Lowry acid

تعتبر الأمونيا NH_3 في التفاعل 2 حمض برونستد - لوري

α

The ammonia NH_3 in reaction 1 is a Lewis acid

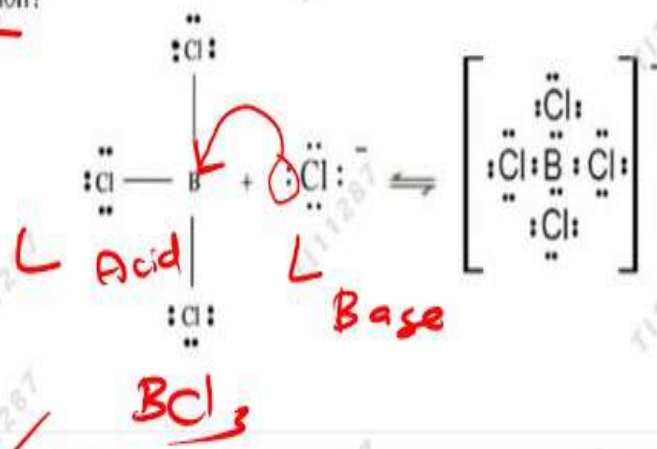
تعتبر الأمونيا NH_3 في التفاعل 1 حمض لويس

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 الخوانيج

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Why does BCl_3 represent Lewis's acid in the following reaction?

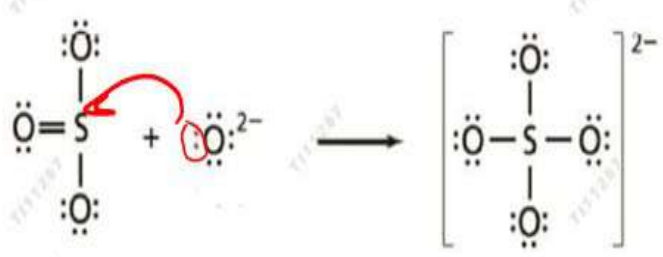
لماذا يُمثل BCl_3 حمض لويس في التفاعل التالي؟



- | | | |
|---|--|---|
| <p>a</p> <p>b</p> <p>c</p> <p>d</p> | <p>~~Because it is proton acceptor from the base Cl^-~~</p> <p>~~Because it is proton donor to the base Cl^-~~</p> <p>Because it is an electron pair donor to the base Cl^-</p> <p><u>Because it is an electron pair acceptor from the base Cl^-</u></p> | <p>لأنه مستقبل للبروتون من القاعدة Cl^-</p> <p>لأنه مانح للبروتون إلى القاعدة Cl^-</p> <p>لأنه مانح لزوج إلكترونات إلى القاعدة Cl^-</p> <p>لأنه مستقبل لزوج إلكترونات من القاعدة Cl^-</p> |
|---|--|---|

Why does O^{2-} represent Lewis's base in the following reaction?

لماذا تُمثل O^{2-} قاعدة لويس في التفاعل التالي؟



- | | |
|--|---|
| <p>~~Because it accepted a proton from SO_3~~</p> <p>~~Because it donated a proton to SO_3~~</p> <p><u>Because it donated a pair of electrons to SO_3</u></p> <p>Because it accepted a pair of electrons from SO_3</p> | <p>لأنها استقبلت بروتوناً من SO_3 ○</p> <p>لأنها منحت بروتوناً إلى SO_3 ○</p> <p>لأنها منحت زوجاً من الإلكترونات إلى SO_3 ○</p> <p>لأنها استقبلت زوجاً من الإلكترونات من SO_3 ○</p> |
|--|---|

What is the substance that contains hydrogen, and ionizes to produce hydrogen ions in aqueous solution?

ما المادة التي تحتوي على هيدروجين وتتأين لإنتاج أيونات الهيدروجين في المحلول المائي؟

① H^+

~~X~~ Lewis acid

حمض لويس

0

~~X~~ Lewis base

قاعدة لويس

0

✓ Arrhenius acid

H^+
produces H^+

حمض أرهينيوس

0

Arrhenius base

OH^-
produces OH^-

قاعدة أرهينيوس

0

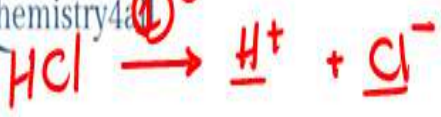
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13	CHM.5.3.04.003.01 Compare between strong and weak acids (using examples, particulate diagrams and ionization equations)
14	CHM.5.3.04.006.01 Define acid ionization constant, K_a , while writing the ionization constant expression for different weak acids
15	CHM.5.3.04.003.02 Relate the strength of weak acids to the numerical values of K_a
16	CHM.5.3.04.003.04 Identify the relationship between the strength of an acid and its conjugate base and the strength of a base and its conjugate acid
17	CHM.5.3.04.006.02 Define base ionization constant, K_b , while writing the ionization constant expression of different weak bases
18	CHM.5.3.04.003.03 Compare between strong and weak bases (using examples, particulate diagrams and ionization equations)

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الخوانيج



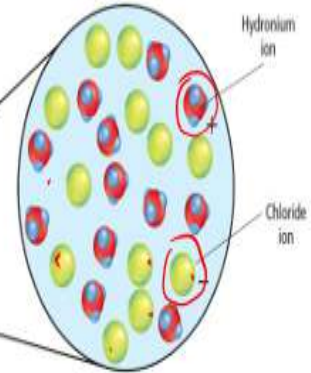
Strengths of Acids



Strong Acid = More ions = High electrical conductivity
in solution



Figure 11 The light glows brightly when electrodes are placed in 0.10M hydrochloric acid solution because all of the HCl is in the form of hydronium ions and chloride ions.



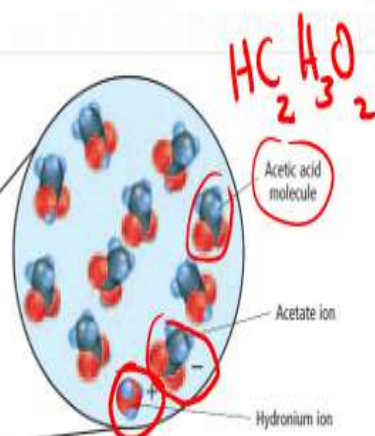
Weak



Weak Acid = Less ions = Low electrical conductivity



Figure 12 When electrodes are placed in 0.10M acetic acid solution, the light is dim. Compare this illustration with Figure 18.11. Explain the difference in the brightness of the bulbs in terms of the concentration of ions in solution.



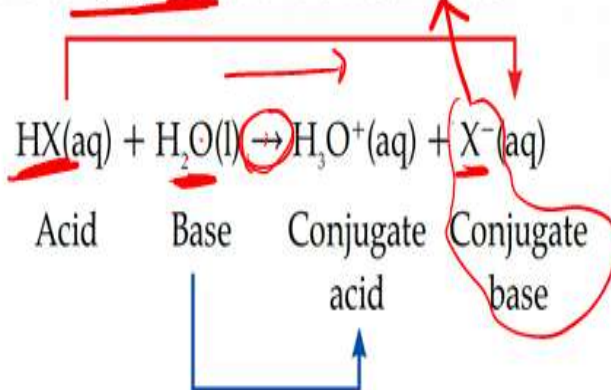
Strengths of Acids

تأين بالكامل

- Acids that ionize completely are called **strong acids**.
- Because they produce the maximum number of hydrogen ions, strong acids are good conductors of electricity.
- Acids that ionize only partially in dilute aqueous solutions are called **weak acids**.



- With a **strong acid**, the conjugate base is a **weak base**.



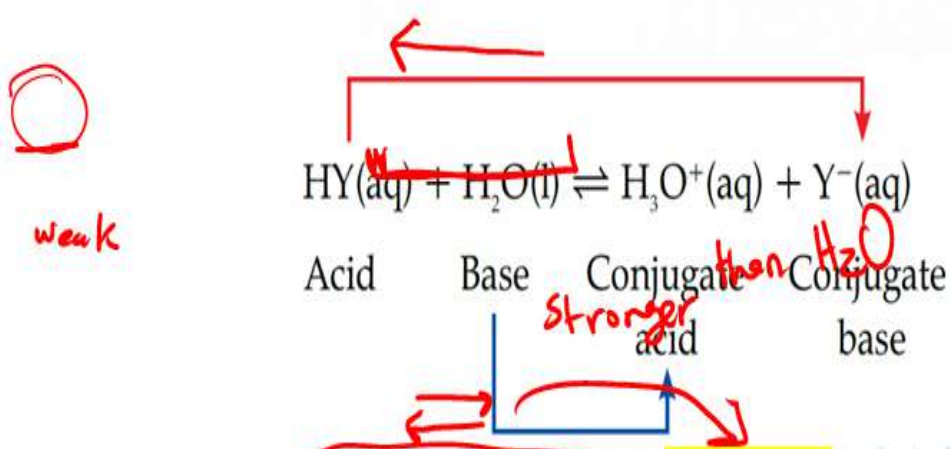
base:- wants gains H^+
 H_2O stronger base X^-

- Equilibrium lies almost completely **to the right**, because the conjugate base has a weaker attraction for the H^+ ion than does the base in the forward reaction.
- Notice that the equation is **shown with a single arrow** to the right for **strong acids**. The reaction goes to completion (right).

H_2O

conjugate base

- In a weak acid, the **ionization equilibrium** lies to the far left because the conjugate base has a greater attraction for H^+ ions than does the base in the forward reaction.



weak

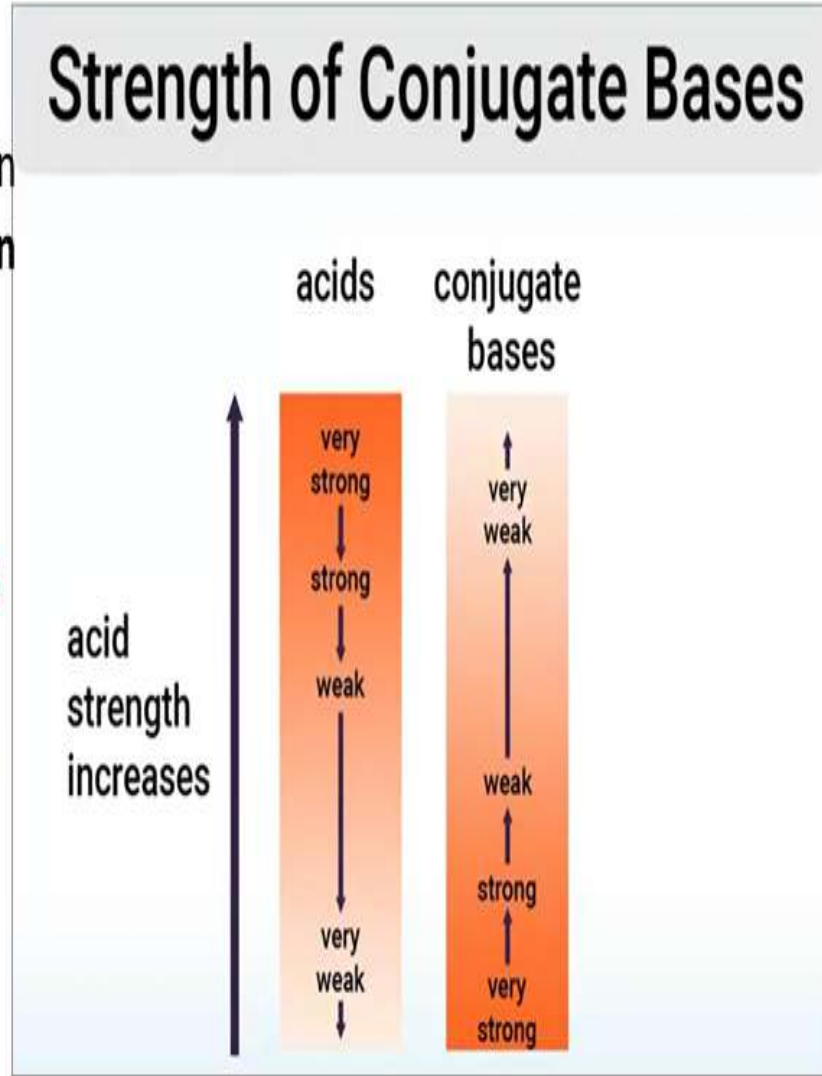
- Notice that the equation is shown with **equilibrium arrows** for **weak acids**.

Strength of Conjugate Bases

What is the relationship between the acid and the conjugate base in terms of strength?

Weak acid = strong conjugate base

Strong acid = weak conjugate base



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Try to Memorize them

CH₃COOH

Table 3 Ionization Equations

Strong Acids		Weak Acids	
Name	Ionization Equation	Name	Ionization Equations
Hydrochloric	$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$	Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$
Hydroiodic	$\text{HI} \rightarrow \text{H}^+ + \text{I}^-$	Acetic	$\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$
Perchloric	$\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$	Hydrosulfuric	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$
Nitric	$\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$	Carbonic	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$
Sulfuric	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$	Hypochlorous	$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$

1. _____ acids dissociate completely in aqueous solutions.

A Only strong

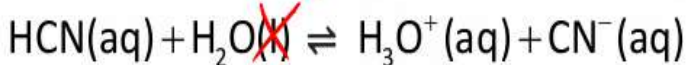
B Only weak

C Both strong and weak

D Neither strong nor weak

Easy chemistry4all

- The equilibrium constant, K_{eq} , provides a measure of the degree of ionization of an acid.



$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$$

- The **acid ionization constant**, K_a , is the value of the equilibrium constant expression for the ionization of a weak acid.

$$K_a = K_{eq} [H_2O] = \frac{[H_3O^+][CN^-]}{[HCN]} = 6.2 \times 10^{-10}$$

- K_a indicates whether products (Ions) or reactants (Acid) are favored (more) at equilibrium.
- For weak acids, the products tend to be smaller compared to the un-ionized molecules (reactant).
- Weaker acids have a smaller K_a . < 1 (More Acid and less ions)

Acid A $K_a = 5$

Acid B $K_a = 3$

A is stronger than B

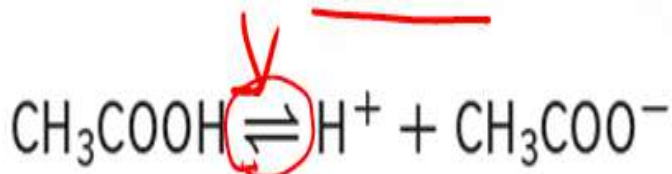
Table 4 Ionization Constants for Weak Acids

Acid	Ionization Equation	K_a (298 K)
Hydrosulfuric, first ionization	$H_2S \rightleftharpoons H^+ + HS^-$	8.9×10^{-8} < 1
Hydrosulfuric, second ionization	$HS^- \rightleftharpoons H^+ + S^{2-}$	1×10^{-19} < 1
Hydrofluoric	$HF \rightleftharpoons H^+ + F^-$	6.3×10^{-4} < 1
Hydrocyanic	$HCN \rightleftharpoons H^+ + CN^-$	6.2×10^{-10} < 1
Acetic	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5} < 1
Carbonic, first ionization	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$	4.5×10^{-7} < 1
Carbonic, second ionization	$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.7×10^{-11} < 1

$K_a > 1$ strong Acid

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2. In the equation below, CH_3COOH is a _____.
 $\text{HC}_2\text{H}_3\text{O}_2$



- A strong acid
- B weak acid**
- C strong base
- D weak base



YouTube

Practice Problems:

APPLICATIONS

12. Write ionization equations and acid ionization constant expressions for each acid.

a. HClO_2

b. HNO_2

c. HIO

13. Write the first and second ionization equations for H_2SeO_3 .

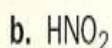
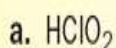
14. **Challenge** Given the expression $K_a = \frac{[\text{AsO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HAsO}_4^{2-}]}$, write the balanced equation for the corresponding reaction.



Practice Problems:

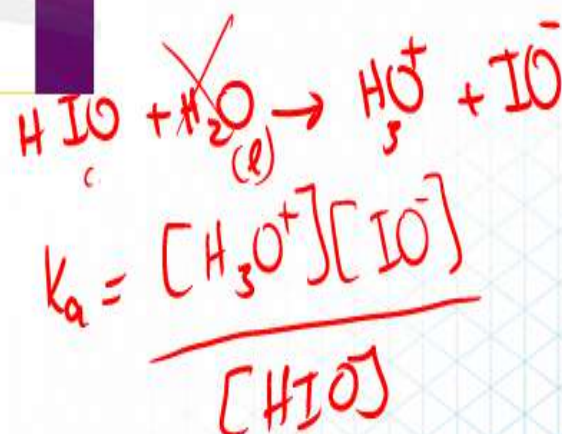
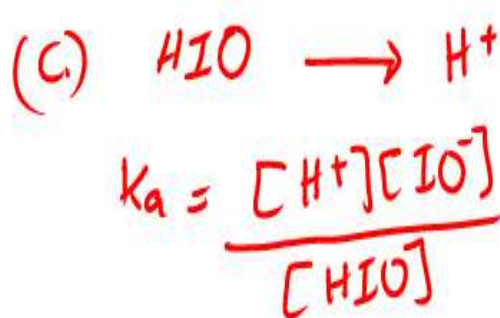
APPLICATIONS

12. Write ionization equations and acid ionization constant expressions for each acid.



13. Write the first and second ionization equations for H_2SeO_3 .

14. **Challenge** Given the expression $K_a = \frac{[\text{AsO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HAsO}_4^{2-}]}$, write the balanced equation for the corresponding reaction.

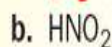
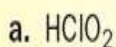


APPLICATIONS

Lesson 1

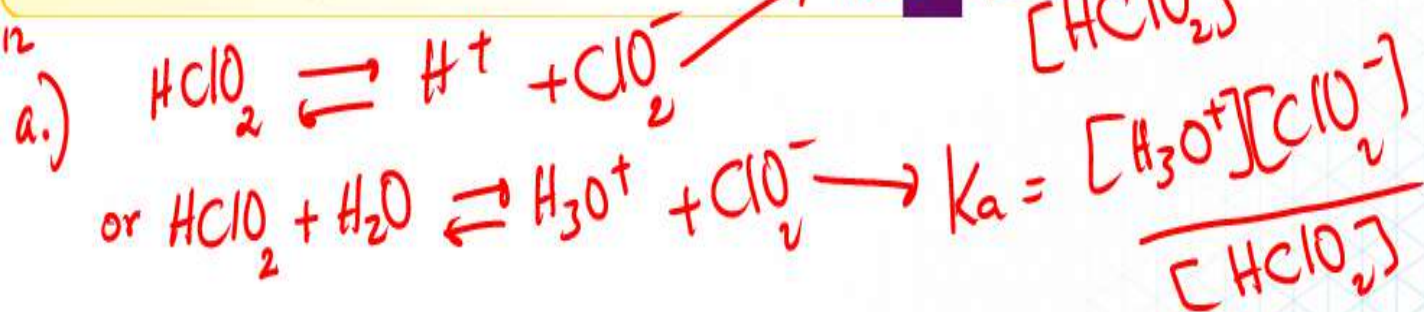
K_a

12. Write ionization equations and acid ionization constant expressions for each acid.



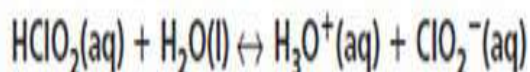
13. Write the first and second ionization equations for H_2SeO_3 .

14. **Challenge** Given the expression $K_a = \frac{[\text{AsO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HAsO}_4^{2-}]}$, write the balanced equation for the corresponding reaction.



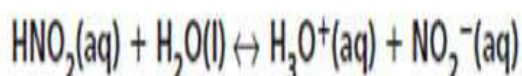
12. Write ionization equations and acid ionization constant expressions for the following acids.

✓ a. HClO_2



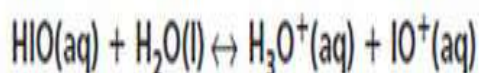
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

✓ b. HNO_2



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

✓ c. HIO



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{IO}^-]}{[\text{HIO}]}$$

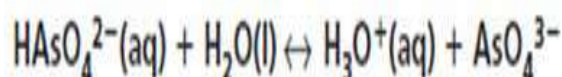
✓ 13. Write the first and second ionization equations for H_2SeO_3 .



✓ 14. Challenge Given the expression

$$K_a = \frac{[\text{AsO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HAsO}_4^{2-}]}, \text{ write the balanced}$$

equation for the corresponding reaction.



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- A base that dissociates completely into metal ions and hydroxide ions is known as a **strong base**.
- A **weak base** ionizes only partially in dilute aqueous solution.

Table 5 **Dissociation Equations for Strong Bases**

metal	$\text{NaOH(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
	$\text{KOH(s)} \rightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$
	$\text{RbOH(s)} \rightarrow \text{Rb}^+(\text{aq}) + \text{OH}^-(\text{aq})$
	$\text{CsOH(s)} \rightarrow \text{Cs}^+(\text{aq}) + \text{OH}^-(\text{aq})$
	$\text{Ca(OH)}_2(\text{s}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$
	$\text{Ba(OH)}_2(\text{s}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$

- The **base ionization constant**, K_b , is the value of the equilibrium constant expression for the ionization of a base.
 - For example, for Methylamine (CH_3NH_2) (Reaction is below in Table 6):
- $$\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$$

$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$$
- The smaller the value of K_b , the weaker the base. $\ll 1$

Table 6 Ionization Constants of Weak Bases

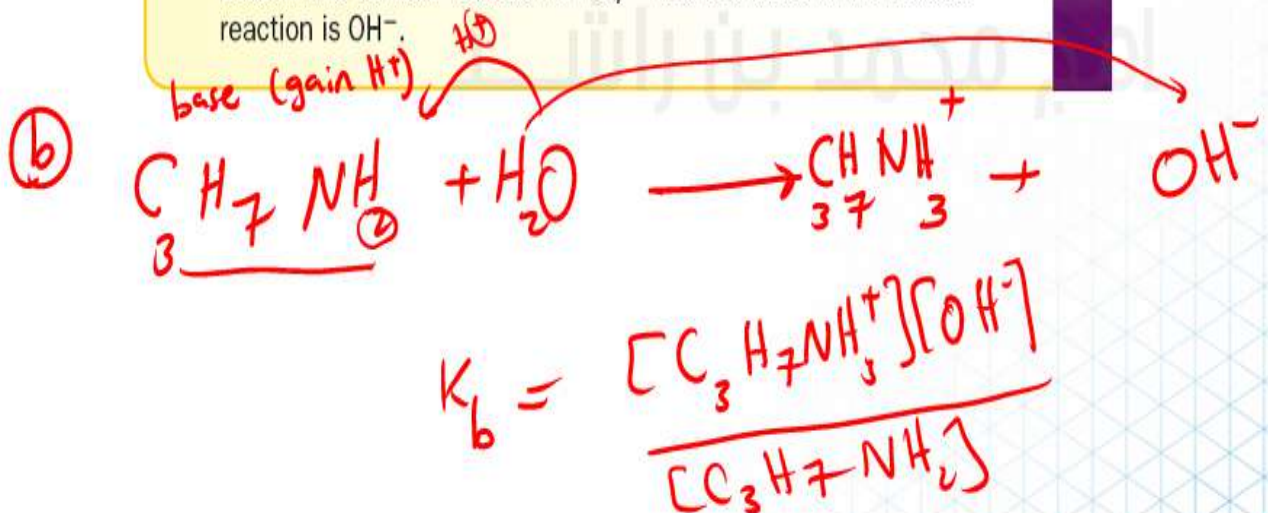
Base	Ionization Equation	K_b (298 K)
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	5.0×10^{-4}
Methylamine	$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	4.3×10^{-4}
Ammonia	$\text{NH}_3(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	2.5×10^{-5}
Aniline	$\text{C}_6\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	4.3×10^{-10}

APPLICATIONS

15. Write ionization equations and base ionization constant expressions for the following bases.

- a. hexylamine ($C_6H_{13}NH_2$) c. carbonate ion (CO_3^{2-})
b. propylamine ($C_3H_7NH_2$) d. hydrogen sulfite ion (HSO_3^-)

16. **Challenge** Write an equation for a base equilibrium in which the base in the forward reaction is PO_4^{3-} and the base in the reverse reaction is OH^- .

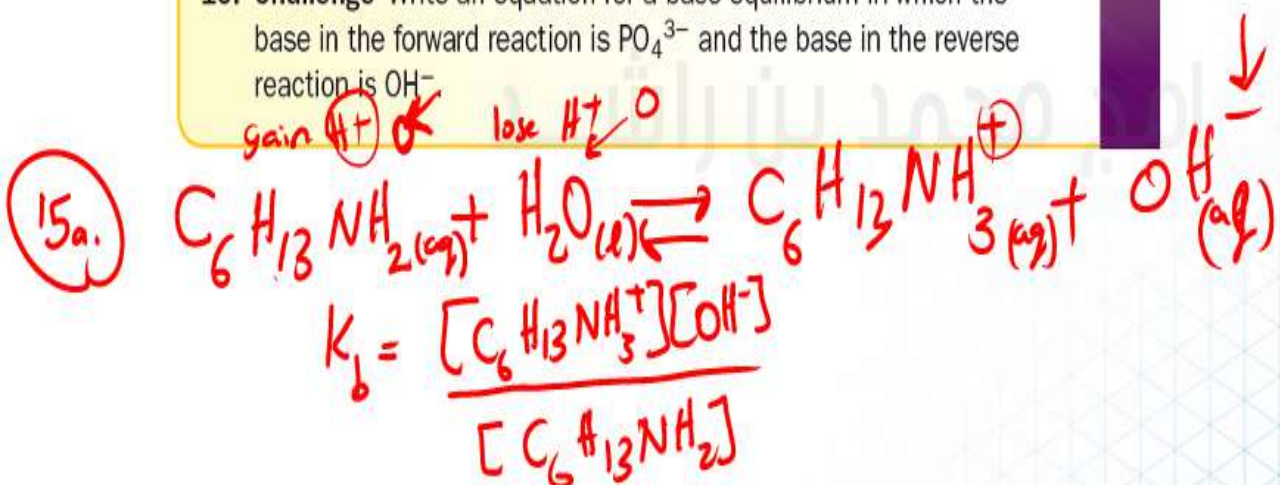


APPLICATIONS

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APPLICATIONS

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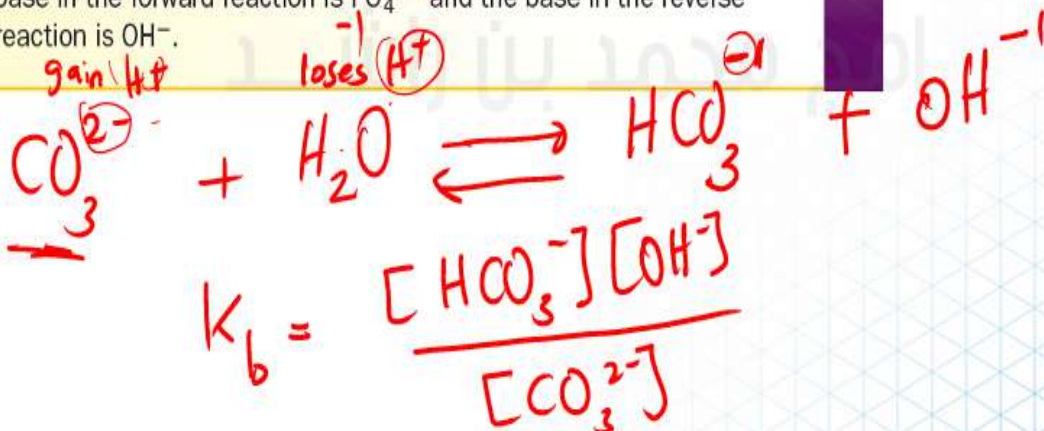
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APPLICATIONS

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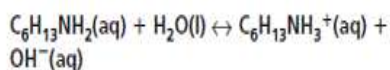
d. hydrogen sulfite ion (HSO_3^-)

16. **Challenge** Write an equation for a base equilibrium in which the base in the forward reaction is PO_4^{3-} and the base in the reverse reaction is OH^- .

asy nemistry4all

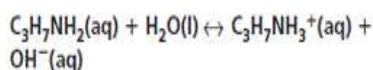
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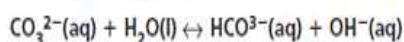
$$K_b = \frac{[C_6H_{13}NH_3^+][OH^-]}{[C_6H_{13}NH_2]}$$

b. propylamine ($C_3H_7NH_2$)



$$K_b = \frac{[C_3H_7NH_3^+][OH^-]}{[C_3H_7NH_2]}$$

c. carbonate ion (CO_3^{2-})



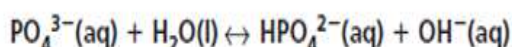
$$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$$

d. hydrogen sulfite ion (HSO_3^-)



$$K_b = \frac{[H_2SO_3][OH^-]}{[HSO_3^-]}$$

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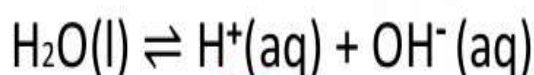


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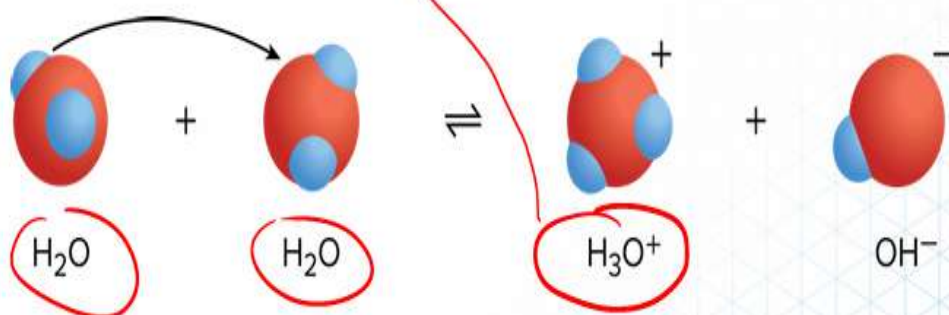
19	CHM.5.3.04.006.03 Relate the acidity and basicity of an aqueous solution to the hydronium and hydroxide ion concentration and pH at 25oC or 298 K
20	CHM.5.3.04.007.01 Use Kw to calculate the hydronium ion and hydroxide ion concentration at a given temperature and vice versa
21	CHM.5.3.04.007.04 Calculate pOH when the $[H^+]$ or $[OH^-]$ is given and vice versa
22	CHM.5.3.04.007.06 Calculate $[H^+]$ and $[OH^-]$ from pH
23	CHM.5.3.04.006.04 Calculate the acid dissociation constant, Ka, given acid concentration, $[H^+]$ and pH
24	CHM.5.3.04.005.07 List different methods used to measure the pH of a solution

Ion Product Constant for Water

- Pure water contains **equal concentrations of H^+ and OH^-** ions produced by self-ionization.
- The equation for the equilibrium can be simplified as follows.

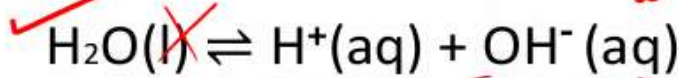


- In the self-ionization of water, one water molecule acts as an acid, and the other acts as a base.



Ion Product Constant for Water

- The **ion product constant for water, K_w** is the value of the equilibrium constant expression for the self-ionization of water.



$$K_w = [H^+][OH^-]$$

- Fact from experiments:** With pure water at 298 K (25 °C), both $[H^+]$ and $[OH^-]$ are equal to $1.0 \times 10^{-7} M$.

$$K_w = [H^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

$$K_w = 1.0 \times 10^{-14}$$

- According to Le Châtelier's Principle, as $[H^+]$ goes up, $[OH^-]$ must go down, and vice versa (والعكس صحيح). **This happens so that the value of K_w will not change.**

22. The concentration of either the H^+ ion or the OH^- ion is given for four aqueous solutions at 298 K. For each solution, calculate $[H^+]$ or $[OH^-]$. State whether the solution is acidic, basic, or neutral.

- a. $[H^+] = 1.0 \times 10^{-13} M$ c. $[OH^-] = 1.0 \times 10^{-3} M$
b. $[OH^-] = 1.0 \times 10^{-7} M$ d. $[H^+] = 4.0 \times 10^{-5} M$

23. **Challenge** Calculate the number of H^+ ions and the number of OH^- ions in 300 mL of pure water at 298 K.

23 d. given $[H^+] = 4 \times 10^{-5} M$
unknown $[OH^-] = ?$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1 \times 10^{-14}}{4 \times 10^{-5}} = 2.5 \times 10^{-10} M$$

$[H^+] > [OH^-]$ acidic

22 The concentration of either the H^+ ion or the OH^- ion is given for four aqueous solutions at 298 K. For each solution, calculate $[H^+]$ or $[OH^-]$. State whether the solution is acidic, basic, or neutral.

a. $[H^+] = 1.0 \times 10^{-13} M$

$$K_w = [H^+][OH^-]$$

$$1.0 \times 10^{-14} = (1.0 \times 10^{-13})[OH^-]$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-13}} = \frac{(1.0 \times 10^{-13})[OH^-]}{1.0 \times 10^{-13}}$$

$$[OH^-] = 1.0 \times 10^{-1} M$$

$[OH^-] > [H^+]$, the solution is basic.

b. $[OH^-] = 1.0 \times 10^{-7} M$

$$K_w = [H^+][OH^-]$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = \frac{[H^+](1.0 \times 10^{-7})}{1.0 \times 10^{-7}}$$

$$[H^+] = 1.0 \times 10^{-7} M$$

$[OH^-] = [H^+]$, the solution is neutral.

c. $[OH^-] = 1.0 \times 10^{-3} M$

$$K_w = [H^+][OH^-]$$

$$1.0 \times 10^{-14} = [H^+](1.0 \times 10^{-3})$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = \frac{[H^+](1.0 \times 10^{-3})}{1.0 \times 10^{-3}}$$

$$[H^+] = 1.0 \times 10^{-11} M$$

$[OH^-] > [H^+]$, the solution is basic.

d. $[H^+] = 4.0 \times 10^{-5} M$

$$K_w = [H^+][OH^-]$$

$$1.0 \times 10^{-14} = (4.0 \times 10^{-5})[OH^-]$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = \frac{(4.0 \times 10^{-5})[OH^-]}{4.0 \times 10^{-5}}$$

$$[OH^-] = 2.5 \times 10^{-10} M$$

$[H^+] > [OH^-]$, the solution is acidic.

23 **Challenge** Calculate the number of H^+ ions and the number of OH^- ions in 300 mL of pure water at 298 K.

At 298 K, $[H^+] = [OH^-] = 1.0 \times 10^{-7} M$

$$\text{Mol } H^+ = \frac{1.0 \times 10^{-7} \text{ mol}}{1L} \times \frac{1L}{1000 \text{ mL}} \times 300 \text{ mL}$$

$$= 3.0 \times 10^{-8} \text{ mol}$$

$$3.0 \times 10^{-8} \text{ mol } H^+ \text{ ions} \times \frac{6.02 \times 10^{23} \text{ } H^+ \text{ ions}}{1 \text{ mol } H^+}$$

$$= 1.8 \times 10^{16} \text{ } H^+ \text{ ions}$$

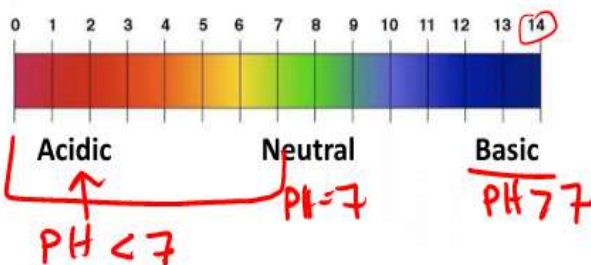
$$\text{Number of } H^+ = \text{number of } OH^-$$

$$= 1.8 \times 10^{16} \text{ ions}$$



pH and pOH

pH Scale

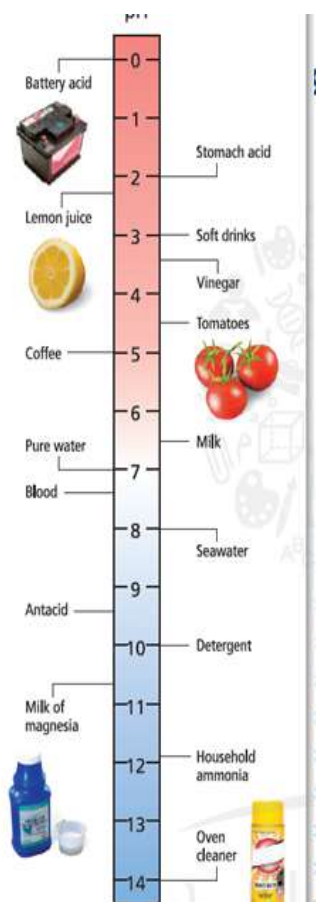


Rules of pH:

When the pH value is closer to zero → It is a strong acid

When the pH value is 7 → It is a neutral compound

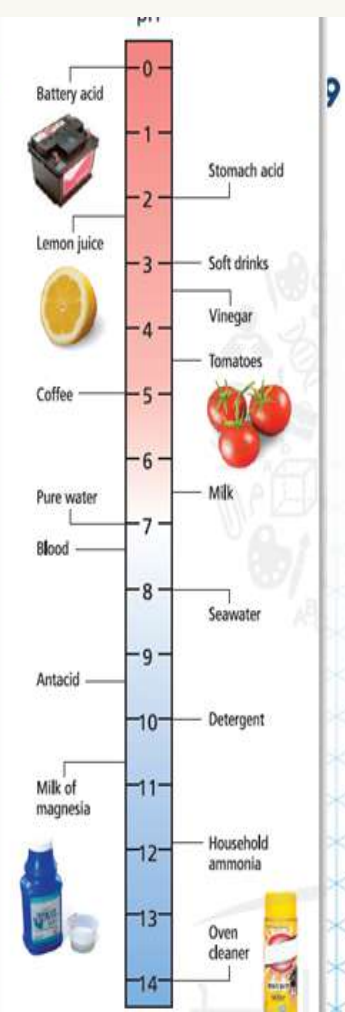
When the pH value is closer to 14 → It is a strong base



pH and pOH

$[H^+]$ $[OH^-]$

- Concentrations of H^+ and OH^- ions are often small numbers expressed in scientific notation.
- pH and pOH are easier ways to express these small concentrations.
- pH** is the negative logarithm of the hydrogen ion concentration of a solution.

$$pH = -\log [H^+]$$


pH and pOH

$[OH^-][H^+] = 1 \times 10^{-14}$

- The **pOH** of a solution is the **negative logarithm** of the **hydroxide ion** concentration.

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

- The sum of pH and pOH is 14.

$$pH + pOH = 14$$

$[H^+]$	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
pOH	14	13	12	11	10	9	8	7	6	5	4	3	2	1	0
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1

← Increasing acidity Neutral Increasing basicity →

Problem types:

مفرد

What you need!!

  hamietrutall
CALCULATE PH FROM $[H^+]$

CALCULATE pOH AND pH FROM $[OH^-]$

CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH

✓ $pH = -\log[H^+]$

✓ $pOH = -\log[OH^-]$

✓ $[OH^-] \times [H^+] = 10^{-14}$

✓ $pH + pOH = 14$

✓ $[H^+] = 10^{-pH}$

✓ $[OH^-] = 10^{-pOH}$

Molarity & pH of strong and weak acids

CALCULATE K_a FROM pH

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APPLICATIONS

24. Calculate the pH of solutions having the following ion concentrations at 298 K.

a. $[H^+] = 1.0 \times 10^{-2}M$ b. $[H^+] = 3.0 \times 10^{-6}M$

25. Calculate the pH of aqueous solutions with the following $[H^+]$ at 298 K.

a. $[H^+] = 0.0055M$ b. $[H^+] = 0.000084M$

26. **Challenge** Calculate the pH of a solution having $[OH^-] = 8.2 \times 10^{-6}M$.

25 (b.) given $[H^+] = 0.000084M$

$$pH = -\log[H^+] = -\log(0.000084) = 4.079$$

APPLICATIONS

24. Calculate the pH of solutions having the following ion concentrations at 298 K.

a. $[H^+] = 1.0 \times 10^{-2}M$ b. $[H^+] = 3.0 \times 10^{-6}M$

25. Calculate the pH of aqueous solutions with the following $[H^+]$ at 298 K.

a. $[H^+] = 0.0055M$ b. $[H^+] = 0.000084M$

26. **Challenge** Calculate the pH of a solution having

$[OH^-] = 8.2 \times 10^{-6}M$.

(1)
 $pOH = -\log[OH^-] = -\log(8.2 \times 10^{-6})$
 $= 5.08$

$pH + pOH = 14$

$pH = 14 - pOH = 14 - 5.08$

$pH = 8.9$

(2)
 $[H^+][OH^-] = 1 \times 10^{-14}$
 $[H^+] = \frac{1 \times 10^{-14}}{[OH^-]} = \frac{1 \times 10^{-14}}{8.2 \times 10^{-6}}$
 $= 1.22 \times 10^{-9}M$

$pH = -\log[H^+] = -\log(1.22 \times 10^{-9})$
 $= 8.9$

24. Calculate the pH of solutions having the following ion concentrations at 298 K.

a. $[H^+] \times 1.0 \times 10^{-2}M$

$$pH = -\log [H^+]$$

$$pH = -\log(1.0 \times 10^{-2})$$

$$pH = 2.00$$

b. $[H^+] = 3.0 \times 10^{-6}M$

$$pH = -\log [H^+]$$

$$pH = -\log(3.0 \times 10^{-6})$$

$$pH = 5.52$$

25. Calculate the pH of aqueous solutions having the following $[H^+]$ at 298 K.

a. $[H^+] = 0.0055M$

$$pH = -\log [H^+]$$

$$pH = -\log 0.0055$$

$$pH = 2.26$$

b. $[H^+] = 0.000084M$

$$pH = -\log [H^+]$$

$$pH = -\log 0.000084$$

$$pH = 4.08$$

26. **Challenge** Calculate the pH of a solution having $[OH^-] = 8.2 \times 10^{-6}M$.

$$[OH^-] = 8.2 \times 10^{-6}M$$

$$K_w = [H^+][OH^-] \times [H^+](8.2 \times 10^{-6})$$

$$[H^+] = \frac{1.0 \times 10^{-14}}{8.2 \times 10^{-6}} = 1.2 \times 10^{-9}$$

$$pH = -\log [H^+]$$

$$pH = -\log(1.2 \times 10^{-9})$$

$$pH = 8.92$$



EXAMPLE 3

CALCULATE pOH AND pH FROM $[OH^-]$ In **Figure 16**, a cow is being fed straw and hay that has been treated with ammonia. The addition of ammonia to animal feed promotes protein growth in the animal. Another use of ammonia is as a household cleaner, which is an aqueous solution of ammonia gas. A typical cleaner has a hydroxide-ion concentration of $4.0 \times 10^{-3}M$. Calculate the pOH and pH of a cleaner at 298 K.

Known

$$[OH^-] = 4.0 \times 10^{-3}M$$

Unknown

$$pOH = ? \quad 2.4$$

$$pH = ?$$

$$pOH = -\log [OH^-] = -\log (4 \times 10^{-3}) = 2.4$$

$$pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 2.4 = 11.6$$

EXAMPLE 3

CALCULATE pOH AND pH FROM $[\text{OH}^-]$ In Figure 16, a cow is being fed straw and hay that has been treated with ammonia. The addition of ammonia to animal feed promotes protein growth in the animal. Another use of ammonia is as a household cleaner, which is an aqueous solution of ammonia gas. A typical cleaner has a hydroxide-ion concentration of $4.0 \times 10^{-3}\text{M}$. Calculate the pOH and pH of a cleaner at 298 K.

Known

$$[\text{OH}^-] = 4.0 \times 10^{-3}\text{M}$$

Unknown

$$\text{pOH} = ?$$

$$\text{pH} = ?$$

2 SOLVE FOR THE UNKNOWN

$$\text{pOH} = -\log [\text{OH}^-]$$

State the equation for pOH.

$$\text{pOH} = -\log (4.0 \times 10^{-3})$$

Substitute $[\text{OH}^-] = 4.0 \times 10^{-3}\text{M}$.

The pOH of the solution is **2.40**.

Use the relationship between pH and pOH to find the pH.

$$\text{pH} + \text{pOH} = 14.00$$

State the equation that relates pH and pOH.

$$\text{pH} = 14.00 - \text{pOH}$$

Solve for pH.

$$\text{pH} = 14.00 - 2.40 = 11.60$$

Substitute pOH = 2.40.

The pH of the solution is **11.60**.

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CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH What are $[H^+]$ and $[OH^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

Known

pH = 7.40

Unknown

$[H^+] = ? \text{ mol/L}$

$[OH^-] = ? \text{ mol/L}$

$[H^+] = 10^{-pH}$
 $[OH^-] = 10^{-pOH}$

$3.98 \sim 4$

①

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

$[OH^-] = \frac{1 \times 10^{-14}}{3.98 \times 10^{-8}} = 2.51 \times 10^{-7} \text{ M}$

②

$pOH = 14 - pH = 14 - 7.4 = 6.6$

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

$[OH^-] = 10^{-pOH} = 10^{-6.6} = 2.51 \times 10^{-7} \text{ M}$

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CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH What are $[H^+]$ and $[OH^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

Known

$$pH = 7.40$$

Unknown

$$[H^+] = ? \text{ mol/L}$$

$$[OH^-] = ? \text{ mol/L}$$

2 SOLVE FOR THE UNKNOWN

Determine $[H^+]$.

$$pH = -\log [H^+]$$

$$-pH = \log [H^+]$$

$$[H^+] = \text{antilog} (-pH)$$

$$[H^+] = \text{antilog} (-7.40)$$

$$[H^+] = 4.0 \times 10^{-8} M$$

3.94

State the equation for pH.

Multiply both sides of the equation by -1 .

Take the antilog of each side to solve for $[H^+]$.

Substitute $pH = 7.40$.

A calculator shows that the antilog of -7.40 is 4.0×10^{-8} .


مجمع زايد التعليمي
الخوانيج

CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH What are $[H^+]$ and $[OH^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

$$[H^+] = 4.0 \times 10^{-8}M$$

A calculator shows that the antilog of -7.40 is 4.0×10^{-8} .

The concentration of H^+ ions in the blood is $4.0 \times 10^{-8}M$.

Determine $[OH^-]$.

$$pH + pOH = 14.00$$

State the equation that relates pH and pOH.

$$pOH = 14.00 - pH$$

Solve for pOH.

$$pOH = 14.00 - 7.40 = 6.60$$

Substitute $pH = 7.40$.

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

State the equation for pOH.

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

Multiply both sides of the equation by -1 .

$$[OH^-] = \text{antilog} (-6.60)$$

Take the antilog of each side and substitute $pOH = 6.60$.

$$[OH^-] = 2.5 \times 10^{-7}M.$$

A calculator shows that the antilog of -6.60 is 2.5×10^{-7} .

The concentration of OH^- ions in the blood is $2.5 \times 10^{-7}M$.

APPLICATIONS

30. Calculate $[H^+]$ and $[OH^-]$ in each of the following solutions.

a. Milk, $pH = 6.50$

b. Lemon juice, $pH = 2.37$

c. Milk of magnesia, $pH = 10.50$

d. Household ammonia, $pH = 11.90$

31. **Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a $pOH = 5.60$.

30
[C.] given $pH = 10.50$ $3.16 \approx 10^0$ $[H^+] =$

unknown $[H^+] = 10^{-pH} = 10^{-10.5} = 3.16 \times 10^{-11} M$

$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{3.16 \times 10^{-11}} = 3.16 \times 10^{-4} M$

$pOH = 14 - 10.5 = 3.5$

$[OH^-] = 10^{-pOH} = 10^{-3.5} = 3.16 \times 10^{-4} M$

APPLICATIONS

30. Calculate $[H^+]$ and $[OH^-]$ in each of the following solutions.

a. Milk, pH = 6.50

b. Lemon juice, pH = 2.37

c. Milk of magnesia, pH = 10.50

d. Household ammonia, pH = 11.90

31. **Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

$$pH = 14 - 5.6 = 8.4$$

$$[H^+] = 10^{-8.4} = 3.98 \times 10^{-9} M$$

$$[OH^-] = 10^{-5.6} = 2.51 \times 10^{-6} M$$

APPLICATIONS

30. Calculate $[H^+]$ and $[OH^-]$ in each of the following solutions.

a. Milk, pH = 6.50

b. Lemon juice, pH = 2.37

c. Milk of magnesia, pH = 10.50

d. Household ammonia, pH = 11.90

31. **Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

30. **Calculate** $[H^+]$ and $[OH^-]$ in each of the following solutions.

a. Milk, pH = 6.50

$$[H^+] = \text{antilog}(-pH)$$

$$[H^+] = \text{antilog}(-6.50) = 3.2 \times 10^{-7} M$$

$$pOH = 14.00 - pH = 14.00 - 6.50 = 7.50$$

$$[OH^-] = \text{antilog}(-pOH)$$

$$[OH^-] = \text{antilog}(-7.50) = 3.2 \times 10^{-8} M$$

b. Lemon juice, pH = 2.37

$$[H^+] = \text{antilog}(-pH)$$

$$[H^+] = \text{antilog}(-2.37) = 4.3 \times 10^{-3} M$$

$$pOH = 14.00 - pH = 14.00 - 2.37 = 11.63$$

$$[OH^-] = \text{antilog}(-pOH)$$

$$[OH^-] = \text{antilog}(-11.63) = 2.3 \times 10^{-12} M$$

c. Milk of magnesia, pH = 10.50

$$[H^+] = \text{antilog}(-pH)$$

$$[H^+] = \text{antilog}(-10.50) = 3.2 \times 10^{-11} M$$

$$pOH = 14.00 - pH = 14.00 - 10.50 = 3.50$$

$$[OH^-] = \text{antilog}(-3.50) = 3.2 \times 10^{-4} M$$

d. Household ammonia, pH 11.90

$$[H^+] = \text{antilog}(-pH)$$

$$[H^+] = \text{antilog}(-11.90) = 1.3 \times 10^{-12} M$$

$$pOH = 14.00 - pH = 14.00 - 11.90 = 2.10$$

$$[OH^-] = \text{antilog}(-2.10) = 7.9 \times 10^{-3} M$$

10^{-pOH}

31. **Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

$$[OH^-] = \text{antilog}(-pOH)$$

$$[OH^-] = \text{antilog}(-5.60) = 2.5 \times 10^{-6} M$$

$$pH = 14.00 - 5.60 = 8.40$$

$$[H^+] = \text{antilog}(-8.40) = 4.0 \times 10^{-9} M$$

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

32. Calculate the K_a for the following acids using the given information.

- a. 0.220M solution of H_3AsO_4 , pH = 1.50 b. 0.0400M solution of $HClO_2$, pH = 1.80

33. Calculate the K_a of the following acids using the given information.

- a. 0.00330M solution of benzoic acid (C_6H_5COOH), pOH = 10.70

- b. 0.100M solution of cyanic acid ($HCNO$), pOH = 11.00

32 c. 0.150M solution of butanoic acid (C_3H_7COOH), pOH = 11.18

Challenge Calculate the K_a of a 0.0091M solution of an unknown acid (HX) having a pOH of 11.32. Use Table 18.4 to identify to acide.

32. Calculate the K_a for the following acids using the given information.

- a. 0.220M solution of H_3AsO_4 , pH = 1.50 b. 0.0400M solution of $HClO_2$, pH = 1.80

33. Calculate the K_a of the following acids using the given information.

- a. 0.00330M solution of benzoic acid (C_6H_5COOH), pOH = 10.70

- b. 0.100M solution of cyanic acid ($HCNO$), pOH = 11.00

32 c. 0.150M solution of butanoic acid (C_3H_7COOH), pOH = 11.18

Challenge Calculate the K_a of a 0.0091M solution of an unknown acid (HX) having a pOH of 11.32. Use Table 18.4 to identify to acide.

31) a) given

$$[C_6H_5COOH]_i = 0.0033M$$

$$pOH = 10.7$$

$$pH = 14 - 10.7 = 3.3$$

$$[H^+] = 10^{-pH} = 10^{-3.3} = 5.01 \times 10^{-4}M$$

$$[C_6H_5COO^-] = [H^+] = 5.01 \times 10^{-4}M$$

$$[C_6H_5COOH]_f = 0.0033 - (5.01 \times 10^{-4}) = 2.8 \times 10^{-3}M$$

unknown: K_a



$$K_a = \frac{[H^+][C_6H_5COO^-]}{[C_6H_5COOH]_{e.f.}}$$

$$\Rightarrow K_a = \frac{(5.01 \times 10^{-4})^2}{2.8 \times 10^{-3}} = 8.96 \times 10^{-5}$$

a. 0.220M solution of H_3AsO_4 , $\text{pH} = 1.50$

$$K_a = \frac{[\text{H}^+][\text{H}_2\text{AsO}_4^-]}{[\text{H}_3\text{AsO}_4]}$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-1.50) = 3.2 \times 10^{-2} \text{M}$$

$$[\text{H}_2\text{AsO}_4^-] = [\text{H}^+] = 3.2 \times 10^{-2} \text{M}$$

$$[\text{H}_3\text{AsO}_4] = 0.220 \text{M} - 3.2 \times 10^{-2} \text{M} = 0.188 \text{M}$$

$$K_a = \frac{(3.2 \times 10^{-2})(3.2 \times 10^{-2})}{0.188} = 5.4 \times 10^{-3}$$

b. 0.0400M solution of HClO_2 , $\text{pH} = 1.80$

$$K_a = \frac{[\text{H}^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-1.80) = 1.6 \times 10^{-2} \text{M}$$

$$[\text{ClO}_2^-] = [\text{H}^+] = 1.6 \times 10^{-2} \text{M}$$

$$[\text{HClO}_2] = 0.0400 \text{M} - 1.6 \times 10^{-2} \text{M} = 0.024 \text{M}$$

$$K_a = \frac{(1.6 \times 10^{-2})(1.6 \times 10^{-2})}{0.024} = 1.1 \times 10^{-2}$$

31. Calculate the K_a of the following acids using the given information.

a. 0.00330M solution of benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$), $\text{pOH} = 10.70$

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 10.70 = 3.30$$

$$[\text{H}^+] = \text{antilog}(-\text{pH}) = 10^{-\text{pH}}$$

$$[\text{H}^+] = \text{antilog}(-3.30) = 5.0 \times 10^{-4} \text{M}$$

$$[\text{C}_6\text{H}_5\text{COO}^-] = [\text{H}^+] = 5.0 \times 10^{-4} \text{M}$$

$$[\text{C}_6\text{H}_5\text{COOH}] = 0.00330 \text{M} - 5.0 \times 10^{-4} \text{M} = 0.0028 \text{M}$$

$$K_a = \frac{[\text{H}^+][\text{C}_6\text{H}_5\text{COO}^-]}{[\text{C}_6\text{H}_5\text{COOH}]}$$

$$= \frac{(5.0 \times 10^{-4})(5.0 \times 10^{-4})}{(0.0028)}$$

$$K_a = 8.9 \times 10^{-5}$$

b. 0.100M solution of cyanic acid (HCNO), $\text{pOH} = 11.00$

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 11.00 = 3.00$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-3.00) = 1.0 \times 10^{-3} \text{M}$$

$$[\text{CNO}^-] = [\text{H}^+] = 1.0 \times 10^{-3} \text{M}$$

$$[\text{HCNO}] = 0.100 - 1.0 \times 10^{-3} \text{M} = 0.099 \text{M}$$

$$K_a = \frac{[\text{H}^+][\text{CNO}^-]}{[\text{HCNO}]} = \frac{(1.0 \times 10^{-3})(1.0 \times 10^{-3})}{(0.099)}$$

$$K_a = 1.0 \times 10^{-5}$$

c. 0.150M solution of butanoic acid ($\text{C}_3\text{H}_7\text{COOH}$), $\text{pOH} = 11.18$

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 11.18 = 2.82$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-2.82) = 1.5 \times 10^{-3} \text{M}$$

$$[\text{C}_3\text{H}_7\text{COO}^-] = [\text{H}^+] = 1.5 \times 10^{-3} \text{M}$$

$$[\text{C}_3\text{H}_7\text{COOH}] = 0.150 \text{M} - 1.5 \times 10^{-3} \text{M} = 0.149 \text{M}$$

$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_7\text{COO}^-]}{[\text{C}_3\text{H}_7\text{COOH}]} = \frac{(1.5 \times 10^{-3})(1.5 \times 10^{-3})}{(0.149)}$$

$$K_a = 1.5 \times 10^{-5}$$

32. **Challenge** Calculate the K_a of a 0.0091M solution of an unknown acid (HX) having a pOH of 11.32. Use **Table 4** to identify the acid.

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 11.32 = 2.68$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-2.68) = 2.1 \times 10^{-3} \text{M}$$

$$[\text{X}^-] = [\text{H}^+] = 2.1 \times 10^{-3} \text{M}$$

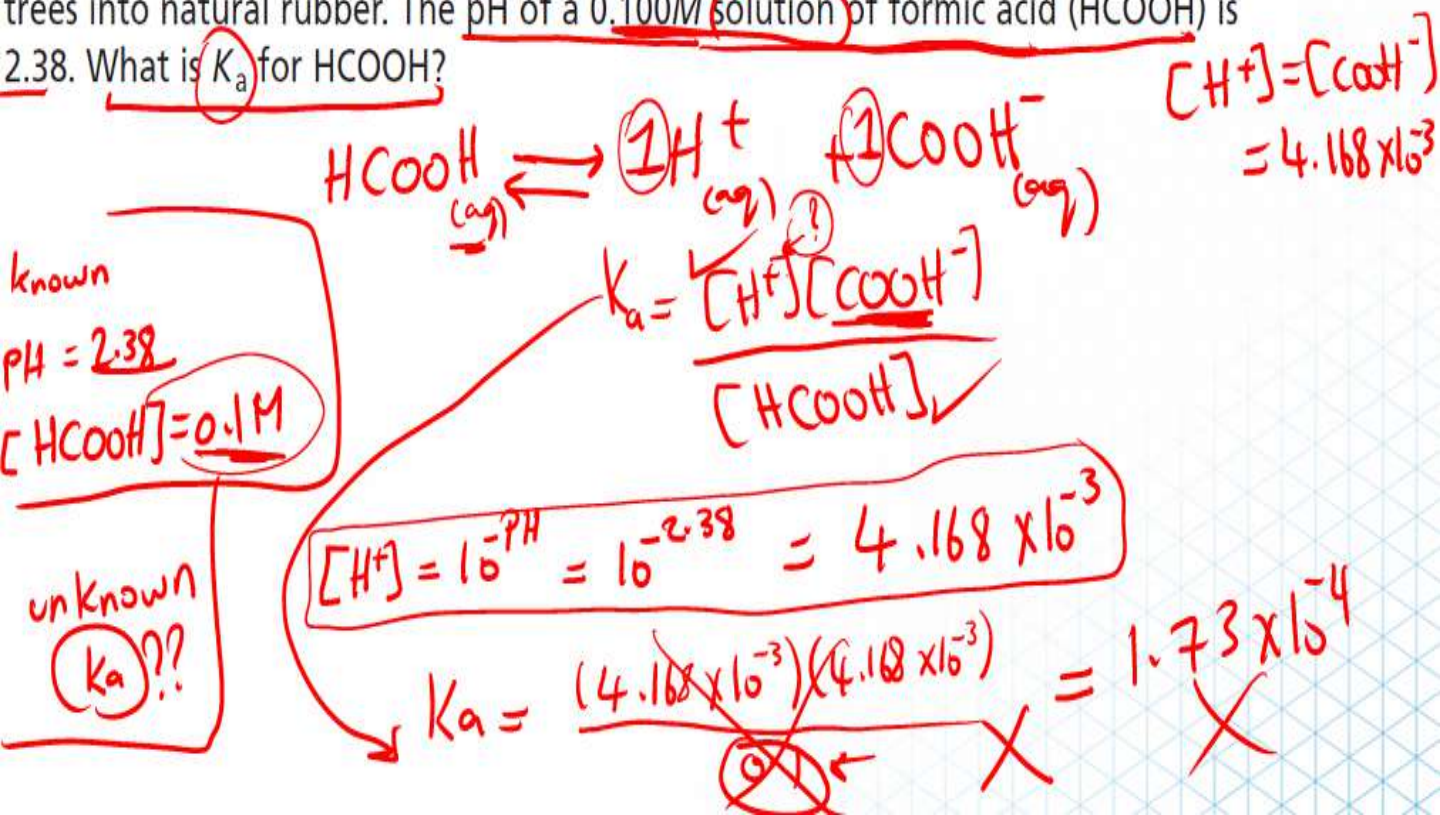
$$[\text{HX}] = 0.0091 - 0.0021 = 0.0070 \text{M}$$

$$K_a = \frac{(0.0021)(0.0021)}{(0.0070)} = 6.3 \times 10^{-4}$$

The acid could be hydrofluoric acid (HF).

EXAMPLE Problem 18.5

Calculate K_a from pH Formic acid is used to process latex tapped from rubber trees into natural rubber. The pH of a 0.100M solution of formic acid (HCOOH) is 2.38. What is K_a for HCOOH ?



Easy Chemistry 4all

EXAMPLE Problem 18.5

Calculate K_a from pH Formic acid is used to process latex tapped from rubber trees into natural rubber. The pH of a 0.100M solution of formic acid (HCOOH) is 2.38. What is K_a for HCOOH ?

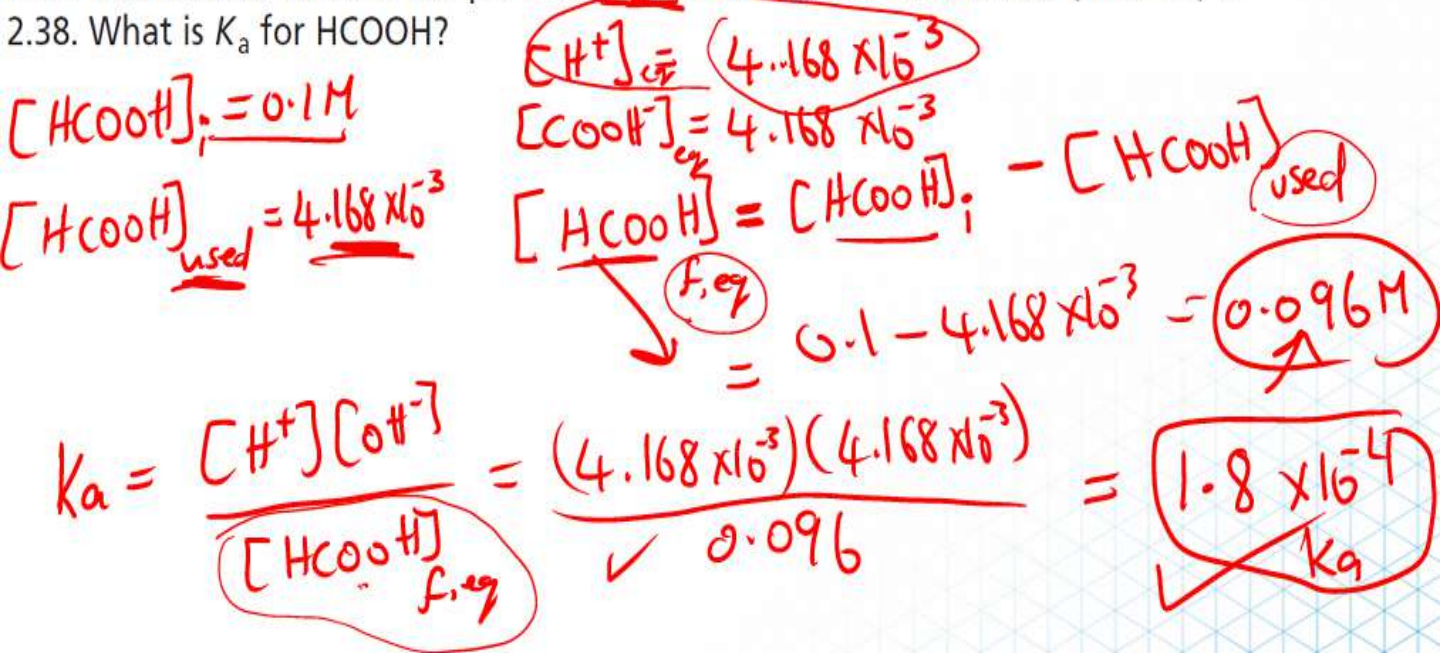




Figure 16 The approximate pH of a solution can be obtained by wetting a piece of pH paper with the solution and comparing the color of the wet paper with a set of standard colors as shown in **a**. The portable pH meter in **b**, which is being used to measure the pH of rain water, provides a more accurate measurement in the form of a digital display of the pH.

Measuring pH

Perhaps you have used indicator paper to measure the pH of a solution. When a strip of pH paper is dipped into an acidic or basic solution, the color of the paper changes. The new color is compared with standard pH colors on a chart, as shown in **Figure 16a**. The pH meter in **Figure 16b** provides a more accurate measure of pH. When electrodes are placed in a solution, the meter gives a direct analog or digital readout of pH.

25	CHM.5.3.04.004.01 Define neutralization reaction while writing the neutralization equation (Complete ionic and net ionic equations)
26	CHM.5.3.04.004.02 Define titration, and titrant CHM.5.3.04.004.03 Explain how to carry out an acid–base titration
27	CHM.5.3.04.004.04 Explain the difference between the equivalence point and the end point of titration process
28	CHM.5.3.04.009.02 Describe the titration curve of a weak acid with a strong base with respect to type of salt formed, pH and nature of solution at equivalence point, indicator used and its color change and volume of titrant needed for changing color of indicator
29	CHM.5.3.04.009.02 Describe the titration curve of a weak acid with a strong base with respect to type of salt formed, pH and nature of solution at equivalence point, indicator used and its color change and volume of titrant needed for changing color of indicator
30	CHM.5.3.04.004.06 Calculate the molarity (concentration) and volume of a solution using titration data

Reactions Between Acids and Bases

- A **neutralization reaction** is a reaction in which an acid and a base in an aqueous solution react to produce a salt and water.
- A **salt** is an ionic compound made up of a **cation(+ ion)** from a base and an **anion (- ion)** from an acid.

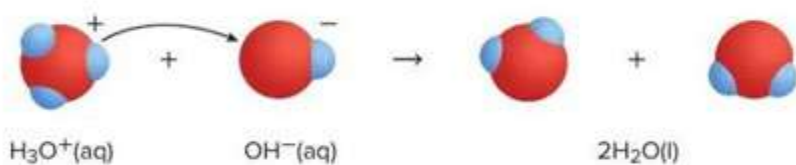
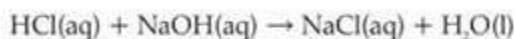
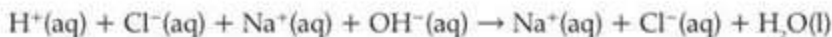


Figure 18 A hydronium ion transfers a hydrogen ion to a hydroxide ion. The loss of the hydrogen ion by H_3O^+ results in a water molecule. The gain of a hydrogen ion by OH^- also results in a water molecule.

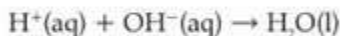
When writing neutralization equations, you must know whether all of the reactants and products in the solution exist as molecules or as formula units. For example, examine the formula equation and complete ionic equation for the reaction between hydrochloric acid and sodium hydroxide.



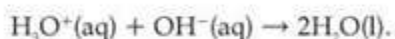
Because HCl is a strong acid, NaOH a strong base, and NaCl a soluble salt, all three compounds exist as ions in an aqueous solution.



The chloride ion and the sodium ion appear on both sides of the equation, so they are spectator ions. They can be eliminated to obtain the net ionic equation for the neutralization of a strong acid by a strong base.

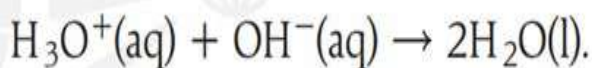


Recall that in an aqueous solution, a H^+ ion exists as a H_3O^+ ion, so the net ionic equation for an acid-base neutralization reaction is

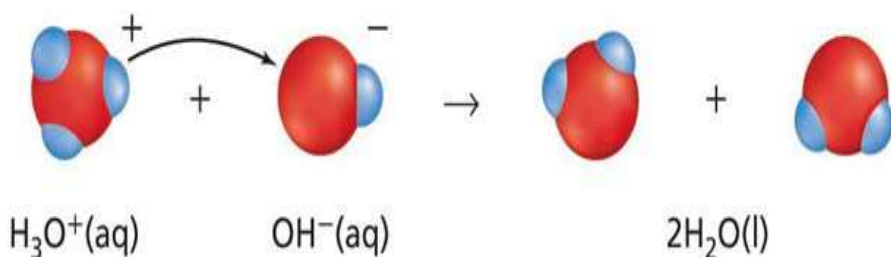


This neutralization reaction is illustrated in **Figure 18**.

Recall that in an aqueous solution, a H^+ ion exists as a H_3O^+ ion, so the net ionic equation for an acid-base neutralization reaction is



■ **Figure 20** A hydronium ion transfers a hydrogen ion to a hydroxide ion. The loss of the hydrogen ion by H_3O^+ results in a water molecule. The gain of a hydrogen ion by OH^- also results in a water molecule.



How can you find its concentration (Molarity)?

Acid-Base Titration





Acid-Base Titration

- **Titration** is a method for determining the concentration of a solution by reacting a known volume of that solution with a solution of known concentration.

Using solution of known
concentration to determine the
concentration of another
solution



*Let's Watch a
video*

- Known concentration
- Unknown concentration

■ **Figure 21** In the titration of an acid by a base, the pH meter measures the pH of the acid solution in the beaker as a solution of a base with a known concentration is added from the buret.



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- **Titration** is the process in which an acid-base neutralization reaction is used to determine the concentration of a solution of unknown concentration.

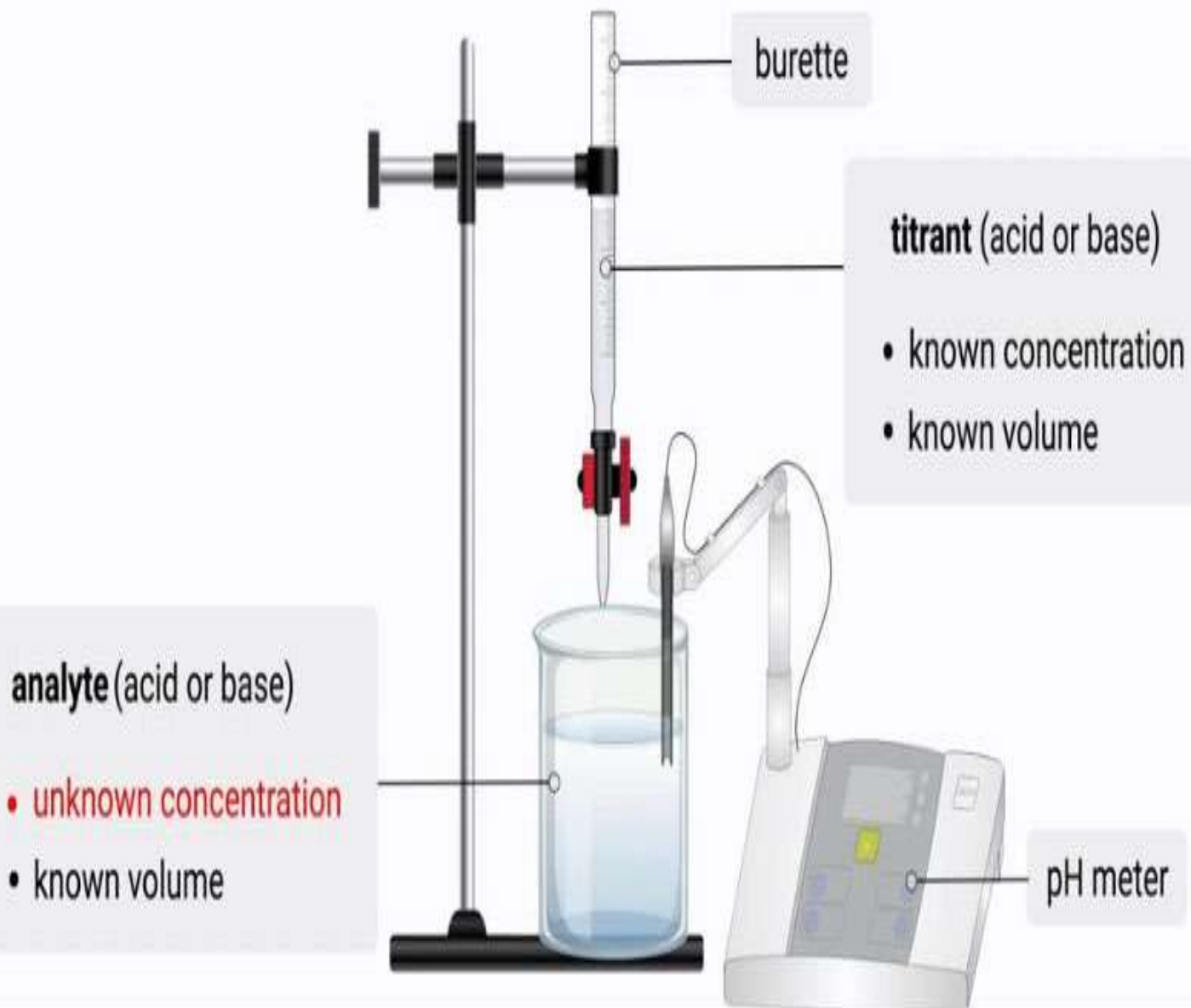
Titration Procedure

How is an acid-base titration performed?

1. A measured volume of an acidic or basic solution of unknown concentration (**called analyte**) is placed in a beaker. The initial pH of the solution is read with a pH meter.
2. A buret is filled with the titrating solution of known concentration. This is the standard solution, or **titrant**.
3. Measured volumes of the standard solution are added slowly to the beaker. The pH is read after each addition. This process continues until the reaction reaches the **equivalence point**, the point at which moles of H^+ ions from the acid equal moles of OH^- ions from the base.

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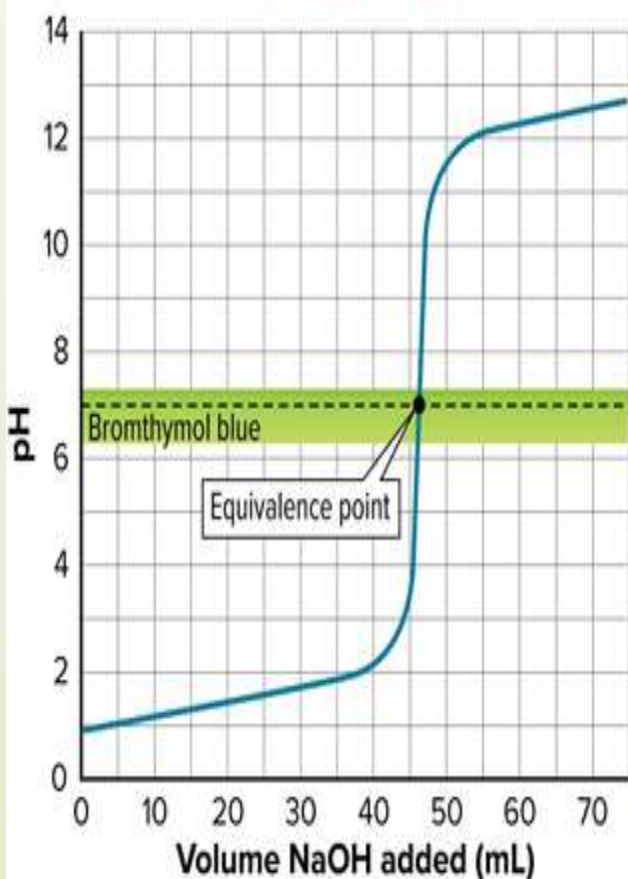
Acid-Base Titrations



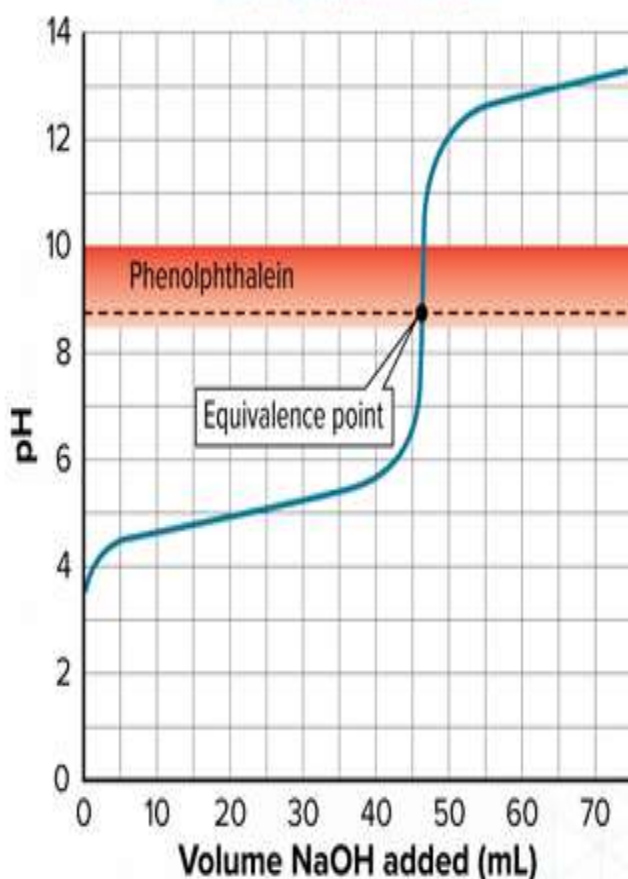
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Acid-Base Titration

50.0 mL 0.100M HCl Titrated
with 0.100M NaOH



50.00 mL 0.1000M HCOOH Titrated
with 0.1000M NaOH

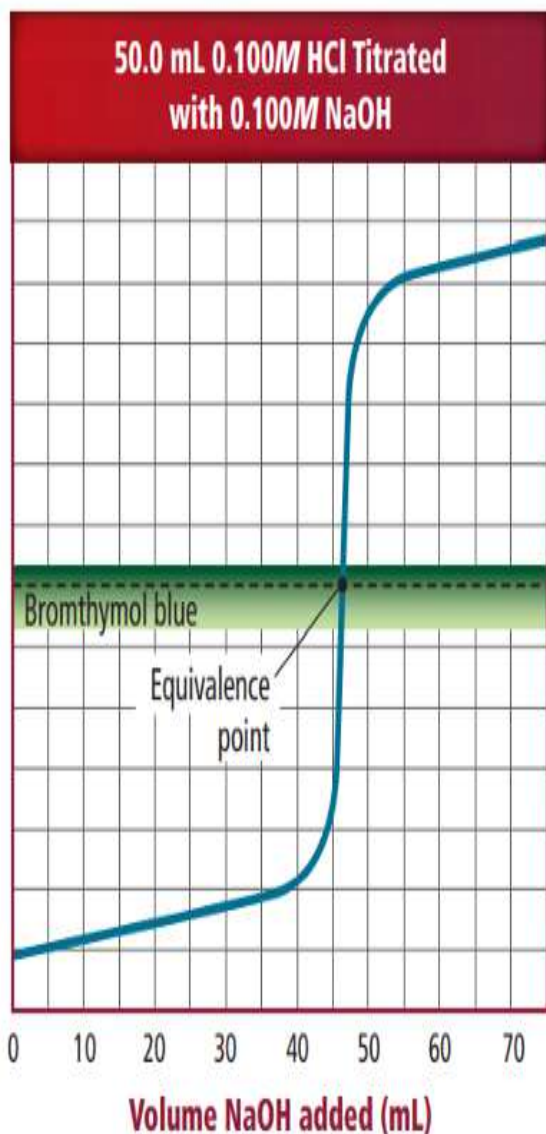


In each titration curve, a steep rise in the pH of the solution indicates that all of the H^+ ions from the acid have been neutralized by the OH^- ions of the base.

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Easy chemistry

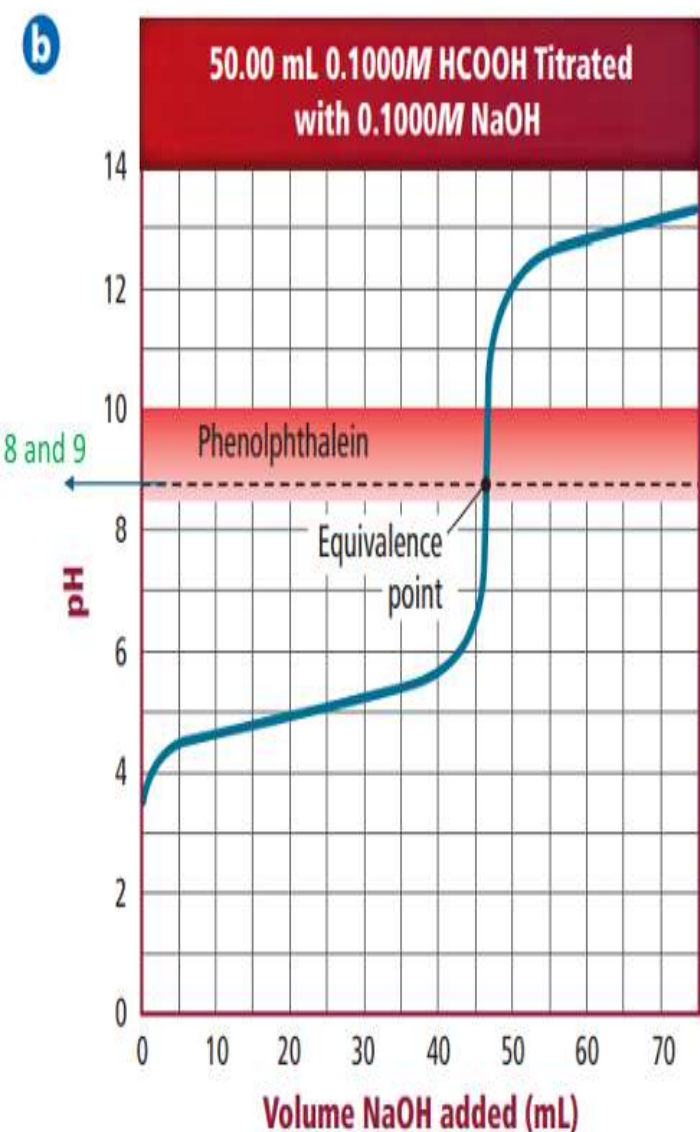
a



strong acid, with 0.100M NaOH, a strong base.

- The initial pH of the 0.100M HCl is 1.00
- As NaOH is added, the acid is neutralized, and the solution's pH increases gradually.
- But, when nearly all the H^+ ions from the acid have been used up, the pH increases dramatically with the addition of an exceedingly small volume of NaOH. Abrupt increase in pH occurs at the equivalence point of the titration.
- Beyond the equivalence point, the addition of more NaOH again results in a gradual increase in PH
- Bromthymol blue is a good choice for a titration of a strong acid with a strong base

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- Titration of weak acid (PH 3.5) with Strong base
- The equivalence point for the titration of Methanoic acid (a weak acid) with sodium hydroxide (a strong base) lies between pH 8 and pH 9
- Phenolphthalein changes color at the equivalence point of a titration of a weak acid with a strong base

2. What term refers to the point in a titration when the concentration of H^+ ions from the acid equals the concentration of OH^- ions from the base?

- ☐ the buffer capacity
- ☐ the titration point
- ☒ the equivalence point **CORRECT**
- ☐ the turning point

The equivalence points of different titrations

At the equivalence point of:

- a weak acid-strong base titration, the pH is greater than 7 ($\text{pH} > 7$)
- a weak base-strong acid titration, the pH is less than 7 ($\text{pH} < 7$)
- **A Strong acid-strong base titration, the $\text{pH} = 7$**

- You might think that all titrations must have an equivalence point at pH 7 because that is the point at which concentrations of hydrogen ions and hydroxide ions are equal and the solution is neutral.
- This is not the case, however. Some titrations have equivalence points at pH values less than 7, and some have equivalence points at pH values greater than 7.
- These differences occur because of reactions between the newly formed salts and water (Later).
- **Equivalence point:** which is the point at which moles of H^+ ion from the acid equal moles of OH^- ion from the base.
- **End point:** The point at which the indicator used in a titration changes color.

Select the **CORRECT** answer to complete the sentence.

The titration process is a method used to _____.

☒ determine the concentration of a solution

☐ determine the volume of a known titrant

☐ measure pH of an acidic or basic solution

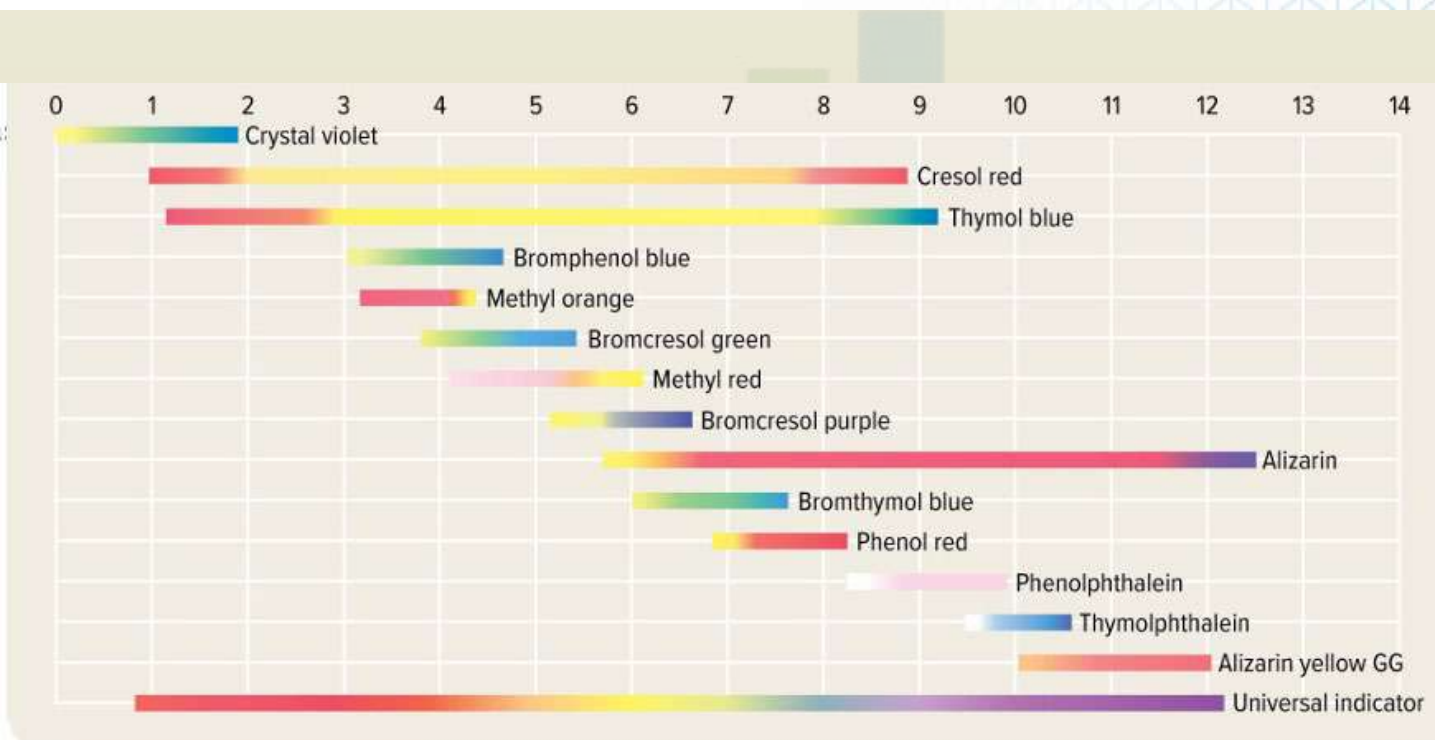
☐ perform acid-base neutralization reactions

How can we know that we reached the equivalence point?

By using an indicator that will change color at the equivalence point

Acid-Base Indicators

- Chemists often use a chemical dye to detect the equivalence point of an acid-base titration.
- Chemical dyes whose colors are affected by acidic and basic solutions (or simply pH) are called **acid-base indicators**.
- The point at which the indicator used in a titration changes color is called the **end point** of the titration.
- It is important to choose an indicator that will change color at the equivalence point of the titration.



Choosing the right indicator is important. The indicator must change color at the equivalence point of the titration which is not always at pH 7.

Tube	Image	Description
1		The buret contains the standard solution (0.1000M NaOH), and the flask contains 25.00 mL HCOOH solution along with a small amount of phenolphthalein indicator.
2		The standard solution is added slowly to the acid solution. The phenolphthalein indicator turns pink, but the color disappears upon mixing, until the end point is reached.
3		The end point of the titration is marked by a permanent, but very light, pink color. A careful reading of the buret reveals that 18.28 mL 0.1000M NaOH has been added.

Indicators and titration end point

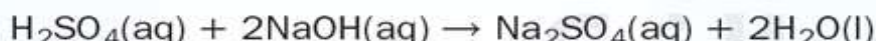
Many indicators used for titration are weak acids. Each has its own particular pH or pH ranges over which it changes color. The point at which the indicator used in a titration changes color is called the **end point** of the titration. It is important to choose an indicator for a titration that will change color at the equivalence point of the titration. Remember that the role of the indicator is to indicate to you, by means of a color change, that just enough of the titrating solution has been added to neutralize the unknown solution.

Figure 23 describes the titration of an unknown solution of methanoic acid (HCOOH) with 0.1000M NaOH.

PROBLEM-SOLVING STRATEGY

Calculating Molarity

The balanced equation for a titration reaction is the key to calculating the unknown molarity. For example, sulfuric acid is titrated with sodium hydroxide according to this equation.



1. Calculate the moles of NaOH in the standard from the titration data: molarity of the base (M_B) and the volume of the base (V_B).
 $(M_B)(V_B) = (\text{mol/L})(\text{L}) = \text{mol NaOH in standard}$
2. From the equation, you know that the mole ratio of NaOH to H_2SO_4 is 2:1. Two moles of NaOH are required to neutralize 1 mol of H_2SO_4 .
 $\text{mol H}_2\text{SO}_4 \text{ titrated} = \text{mol NaOH in standard} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}$
3. M_A represents the molarity of the acid and V_A represents the volume of the acid in liters. $M_A = \frac{\text{mol H}_2\text{SO}_4 \text{ titrated}}{V_A}$

Apply this strategy as you study Example Problem 6.

Molarity from Titration Data A volume of 18.28 mL of a standard solution of 0.1000M NaOH was required to neutralize 25.00 mL of a solution of methanoic acid (HCOOH). What is the molarity of the methanoic acid solution?

Solving Algebraic
 Equations
 pages 954–955

1 Analyze the Problem

You are given the molarity and volume of the NaOH solution and the volume of the methanoic acid (HCOOH) solution. The volume of base used is about four-fifths of the volume of the acid, so the molarity of the acid solution should be less than 0.1M.

Known

- ✓ $V_A = 25.00 \text{ mL HCOOH}$
- ✓ $V_B = 18.28 \text{ mL NaOH}$
- $M_B = 0.1000M$

Unknown

$$M_A = ? \text{ mol/L}$$



$$M_A V_A = M_B \times V_B$$

$$M_A = \frac{M_B \times V_B}{V_A} = \frac{0.1000 \times 18.28}{25.00} = 0.07312 \text{ M}$$

Molarity from Titration Data A volume of 18.28 mL of a standard solution of 0.1000M NaOH was required to neutralize 25.00 mL of a solution of methanoic acid (HCOOH). What is the molarity of the methanoic acid solution?

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Known

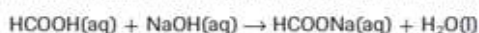
- $V_A = 25.00 \text{ mL HCOOH}$
- $V_B = 18.28 \text{ mL NaOH}$
- $M_B = 0.1000M$

Unknown

$$M_A = ? \text{ mol/L}$$

2 Solve for the Unknown

Write the balanced formula equation for the neutralization reaction.



1 mol NaOH neutralizes 1 mol HCOOH.

$$V_B = 18.28 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.01828 \text{ L}$$

Calculate moles of NaOH.

$$\text{Mol NaOH} = (M_B)(V_B)$$

$$\text{Mol NaOH} = (0.1000 \text{ mol/L})(0.01828 \text{ L}) = 1.828 \times 10^{-3} \text{ mol NaOH}$$

Calculate moles of HCOOH.

$$1.828 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} = 1.828 \times 10^{-3} \text{ mol HCOOH}$$

Calculate the molarity of HCOOH.

$$1.828 \times 10^{-3} \text{ mol HCOOH} = (M_A)(V_A)$$

$$M_A = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{V_A}$$

$$V_A = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.02500 \text{ L HCOOH}$$

$$M_A = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{0.02500 \text{ L HCOOH}} = 7.312 \times 10^{-2} \text{ mol/L}$$

Write the acid to base mole relationship.

Convert volume of base from mL to L.

Apply the relationship between moles of base, molarity of base, and volume of base.

Substitute $M_B = 0.1000M$ and $V_B = 0.01828 \text{ L}$.

Apply the stoichiometric relationship.

Apply the relationship between moles of acids, molarity of acid, and volume of acid.

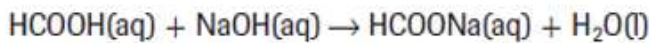
Solve for M_A .

Convert volume of acid from mL to L.

Substitute $V_A = 0.02500 \text{ L}$.

2 Solve for the Unknown

Write the balanced formula equation for the neutralization reaction.



1 mol NaOH neutralizes 1 mol HCOOH.

$$V_B = 18.28 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.01828 \text{ L}$$

Calculate moles of NaOH.

$$\text{Mol NaOH} = (M_B)(V_B)$$

$$\begin{aligned} \text{Mol NaOH} &= (0.1000 \text{ mol/L})(0.01828 \text{ L}) \\ &= 1.828 \times 10^{-3} \text{ mol NaOH} \end{aligned}$$

Calculate moles of HCOOH.

$$\begin{aligned} 1.828 \times 10^{-3} \text{ mol NaOH} &\times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} \\ &= 1.828 \times 10^{-3} \text{ mol HCOOH} \end{aligned}$$

Calculate the molarity of HCOOH.

$$1.828 \times 10^{-3} \text{ mol HCOOH} = (M_A)(V_A)$$

$$M_A = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{V_A}$$

$$V_A = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.02500 \text{ L HCOOH}$$

$$M_A = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{0.02500 \text{ L HCOOH}} = 7.312 \times 10^{-2} \text{ mol/L}$$

Write the acid to base mole relationship.

Convert volume of base from mL to L.

Apply the relationship between moles of base, molarity of base, and volume of base.

Substitute $M_B = 0.1000M$ and $V_B = 0.01828 \text{ L}$.

Apply the stoichiometric relationship.

Apply the relationship between moles of acids, molarity of acid, and volume of acid.

Solve for M_A .

Convert volume of acid from mL to L.

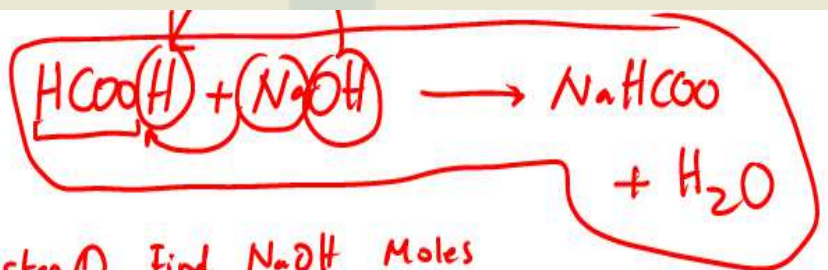
Substitute $V_A = 0.02500 \text{ L}$.

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Use with Example Problem 6.

Problem

A volume of 18.28 mL of a standard solution of 0.1000M NaOH was required to neutralize 25.00 mL of a solution of methanoic acid (HCOOH). What is the molarity of the methanoic acid solution?

titrant
↓



step ① Find NaOH Moles

$$\begin{aligned} \text{Moles} &= M \times \text{Volume} = 0.1 \frac{\text{mol}}{\text{L}} \times \frac{18.28}{1000} \text{ L} \\ &= 1.828 \times 10^{-3} \text{ mol NaOH} \end{aligned}$$

step ② find Moles of HCOOH from Moles of NaOH

Balanced ✓



$$\text{Moles HCOOH} = 1.828 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} = 1.828 \times 10^{-3} \text{ mol HCOOH}$$

$[\text{NaOH}] = 0.1M$
 $M_{\text{NaOH}} = 0.1 \frac{\text{mol}}{\text{L}}$
 $V_{\text{NaOH}} = 18.28 \text{ mL}$
 $V_{\text{HCOOH}} = 25 \text{ mL}$
 $[\text{HCOOH}] = ??$

Titrant
 Standard
 solution
 ↓
 1000 L
 Balanced ✓

Problem

A volume of 18.28 mL of a standard solution of 0.1000M NaOH was required to neutralize 25.00 mL of a solution of methanoic acid (HCOOH). What is the molarity of the methanoic acid solution?

step 3)

$$\text{Molarity} = \frac{\text{Moles}}{\text{Volume (L)}}$$

$$V_{\text{HCOOH}} = 25 \text{ mL} \xrightarrow{\div 1000} \boxed{0.025 \text{ L}}$$

$$\text{Molarity} = [\text{HCOOH}] = \frac{1.828 \times 10^{-3}}{0.025} = 0.07312$$

$$\text{M or } \frac{\text{mol}}{\text{L}}$$

 Or

$$7.312 \times 10^{-2} \text{ mol/L}$$

41 What is the molarity of a nitric acid solution if 43.33 mL of 0.1000M KOH solution is needed to neutralize 20.00 mL of the acid solution?

42 What is the concentration of a household ammonia cleaning solution if 49.90 mL of 0.5900M HCl is required to neutralize 25.00 mL of the solution?

43 **Challenge** How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H₃PO₄?

42

$$M_A V_A = M_B V_B$$

$$M_B = \frac{M_A \times V_A}{V_B} = \frac{49.9 \times 0.59}{25}$$

$$= \boxed{1.17 \text{ M}}$$

$$V_B = 25 \text{ mL}$$

$$V_A = 49.9 \text{ mL}$$

$$M_A = 0.59 \text{ M}$$

Calculation
 ↓ part of Titration

PRACTICE Problems

ADDITIONAL PRACTICE

Concentration $[HNO_3] = ?$

V_{KOH} ↓ $[KOH]$

V_{HNO_3}

41. What is the molarity of a nitric acid solution if 43.33 mL of 0.1000M KOH solution is needed to neutralize 20.00 mL of the acid solution?

42. What is the concentration of a household ammonia cleaning solution if 49.90 mL of 0.5900M HCl is required to neutralize 25.00 mL of the solution?

43. CHALLENGE How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H_3PO_4 ?

$$M_A V_A = M_B V_B \quad (\text{only when there is 1 H \& 1 OH in the formula})$$

$$M_A \times 20 = 0.1 \times 43.33$$

$$M_A = 0.216 M$$

$M \Rightarrow$ concentration or Molarity

$V \Rightarrow$ volume
 $A \rightarrow$ Acid
 $B \rightarrow$ Base

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PRACTICE Problems

ADDITIONAL PRACTICE

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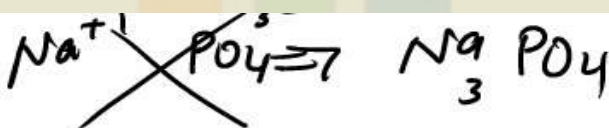
43. CHALLENGE How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H₃PO₄?

42.

$$M_A V_A = M_B V_B$$

$$0.59 \times 49.9 = M_B \times 25$$

$$M_B = 1.17 M$$



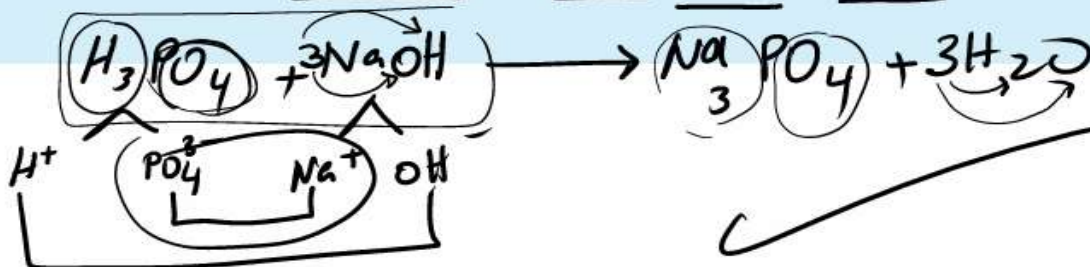
PRACTICE Problems

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$$\text{Molarity} = \frac{\text{mol}}{V(L)}$$

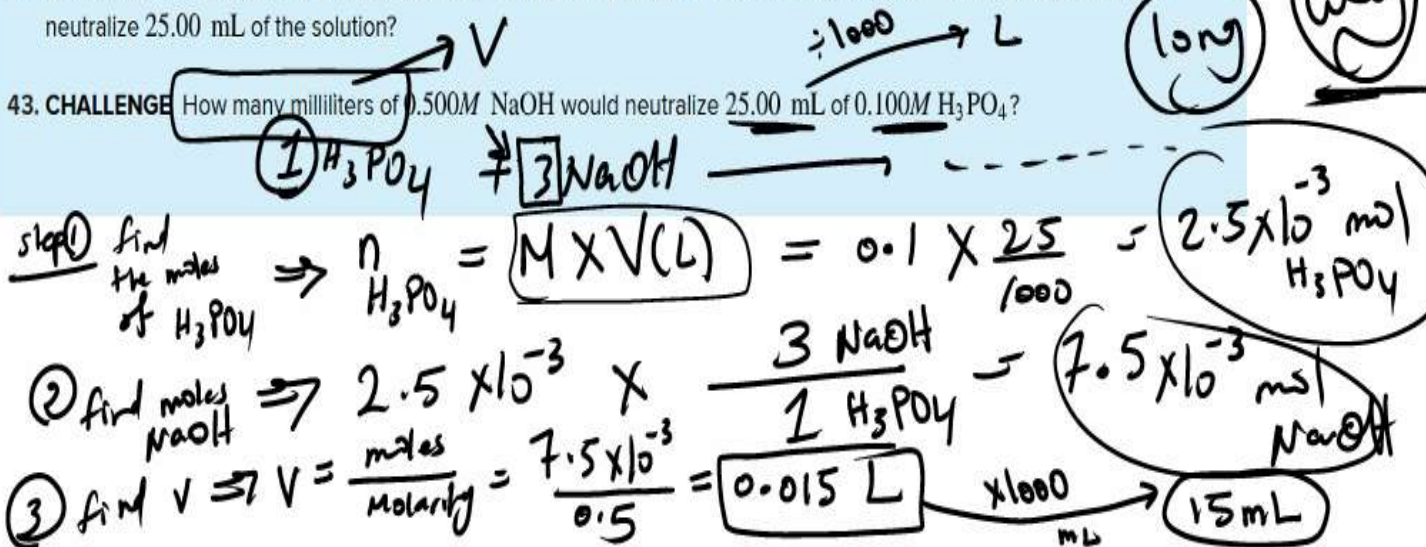
PRACTICE Problems

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43. CHALLENGE How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H_3PO_4 ?



$$\text{H number} \times M_A V_A = M_B V_B \times \text{OH number}$$

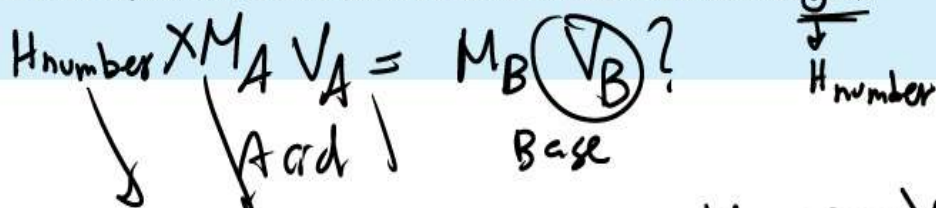
PRACTICE Problems

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43. CHALLENGE How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H_3PO_4 ?



$$3 \times 0.1 \times 25 = 0.5 \times V_B \Rightarrow V_B = 15 \text{ mL}$$

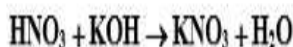
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43. **CHALLENGE** How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H_3PO_4 ?



41. What is the molarity of a nitric acid solution if 43.33 mL 0.1000M KOH solution is needed to neutralize 20.00 mL of the acid solution?



$$43.33 \text{ mL KOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1000 \text{ mol KOH}}{1 \text{ L KOH}}$$

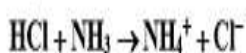
$$= 4.333 \times 10^{-3} \text{ mol KOH}$$

$$4.333 \times 10^{-3} \text{ mol KOH} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol KOH}}$$

$$= 4.333 \times 10^{-3} \text{ mol HNO}_3$$

$$M_{\text{HNO}_3} = \frac{4.333 \times 10^{-3} \text{ mol HNO}_3}{0.02000 \text{ L HNO}_3} = 0.2167 \text{ M}$$

42. What is the concentration of a household ammoniac cleaning solution if 49.90 mL of 0.5900M HCl is required to neutralize 25.00 mL of the solution?



$$49.90 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.5900 \text{ mol HCl}}{1 \text{ L HCl}}$$

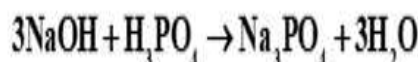
$$= 2.944 \times 10^{-2} \text{ mol HCl}$$

$$2.944 \times 10^{-2} \text{ mol HCl} \times \frac{1 \text{ mol NH}_3}{1 \text{ mol HCl}}$$

$$= 2.944 \times 10^{-2} \text{ mol NH}_3$$

$$M_{\text{NH}_3} = \frac{2.94 \times 10^{-2} \text{ mol NH}_3}{0.02500 \text{ L NH}_3} = 1.178 \text{ M}$$

43. **Challenge** How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H₃PO₄?



$$\text{Mol H}^+ = \frac{0.100 \text{ mol H}_3\text{PO}_4}{1 \text{ L}} \times \frac{3 \text{ mol H}^+}{1 \text{ mol H}_3\text{PO}_4}$$

$$\times 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00750 \text{ mol}$$

$$\text{mol H}^+ = \text{mol OH}^- = 0.00750 \text{ mol}$$

$$(L_{\text{NaOH}})(M_{\text{NaOH}}) = 0.00750 \text{ mol}$$

$$(L_{\text{NaOH}}) = \frac{(0.00750 \text{ mol})}{(0.500 \text{ mol/L})} = 0.0150 \text{ L}$$

$$0.0150 \text{ liter} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 15.0 \text{ mL NaOH}$$

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