

WRITTEN ASSIGNMENT 5

1. Acids

- (i) They have a sour taste.
- (ii) They react with metals to form hydrogen gas and react with bases to form salts and water.
- (iii) They contain hydrogen ions (H^+) when dissolved in water.
- (iv) The presence of an acid turns blue litmus paper to red.

Bases

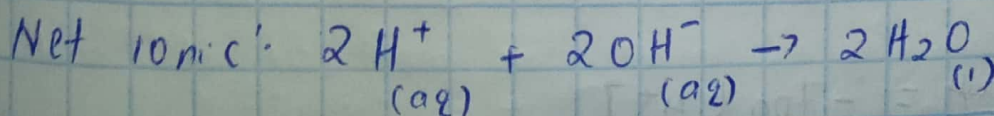
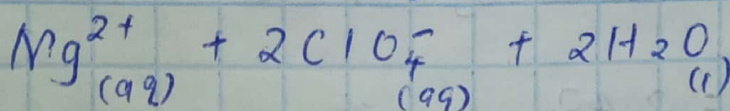
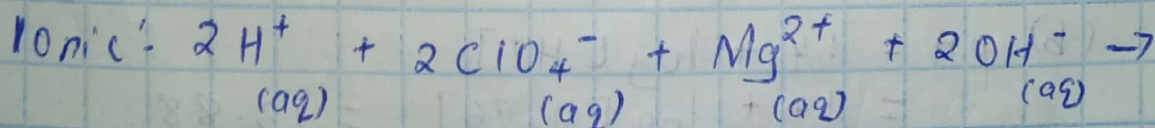
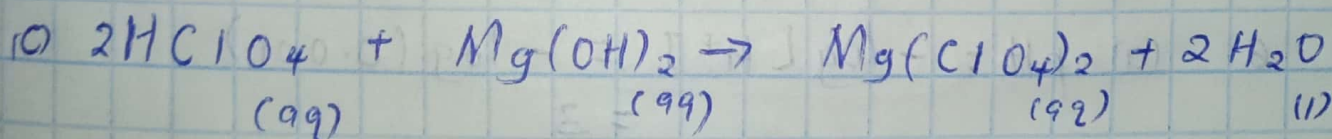
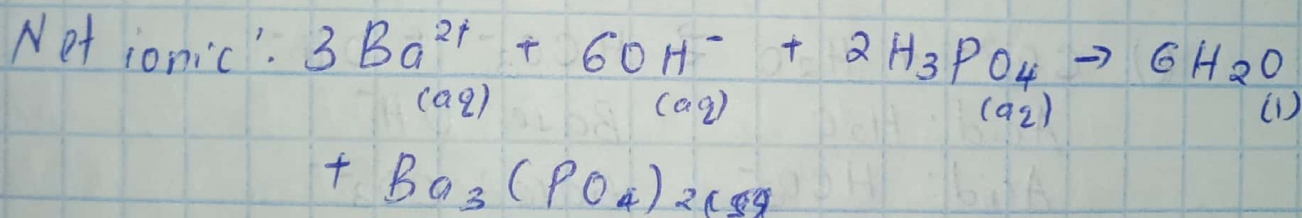
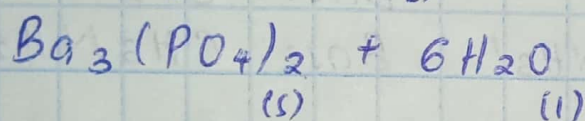
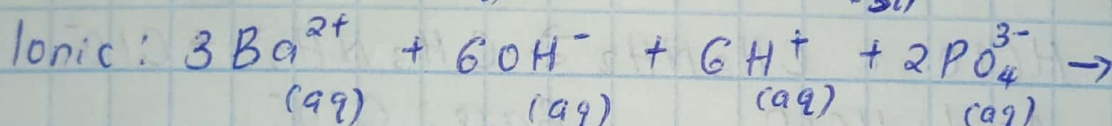
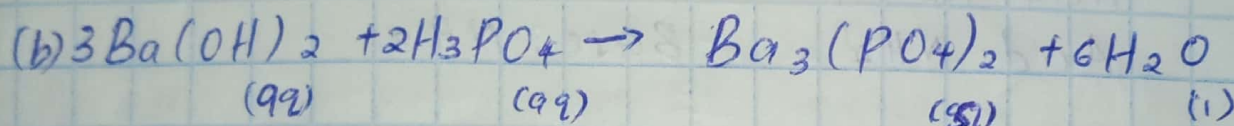
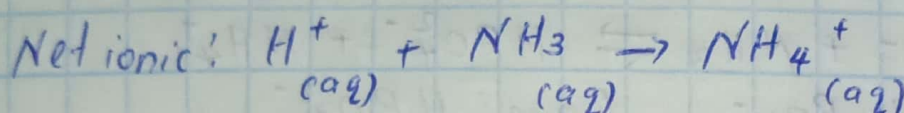
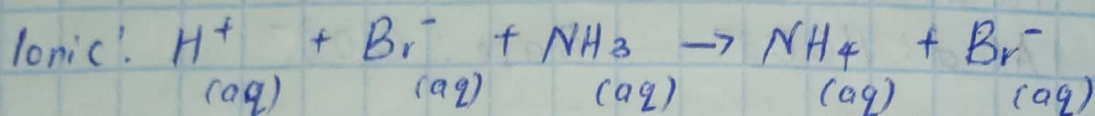
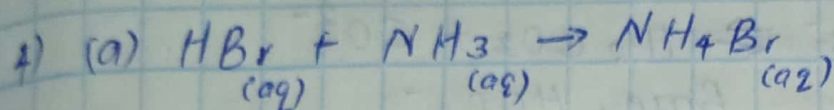
- (i) They have a bitter taste.
- (ii) The presence of a base turns litmus to blue and phenolphthalein to pink.
- (iii) They are slippery to touch (soapy).
- (iv) They do not react with metals.
- (v) They react with acids to form salts and water.

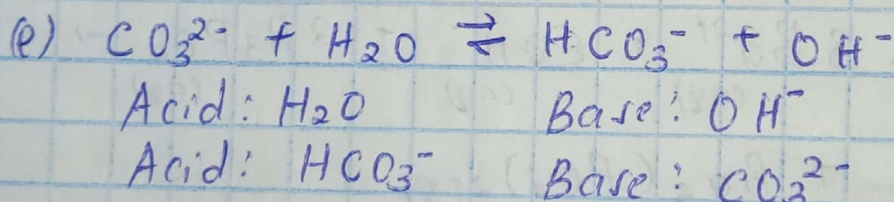
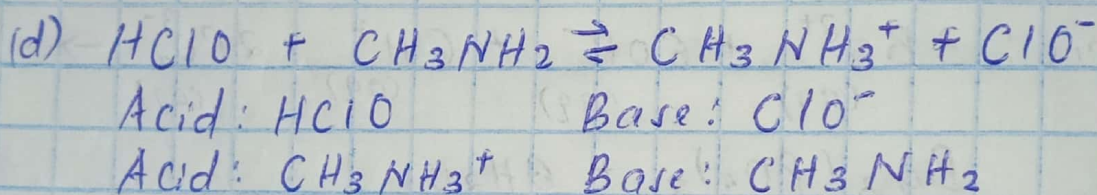
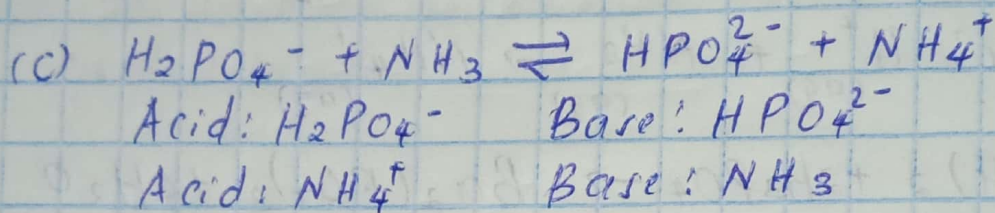
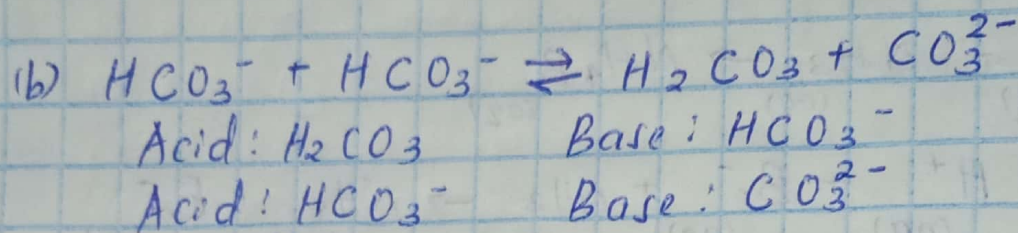
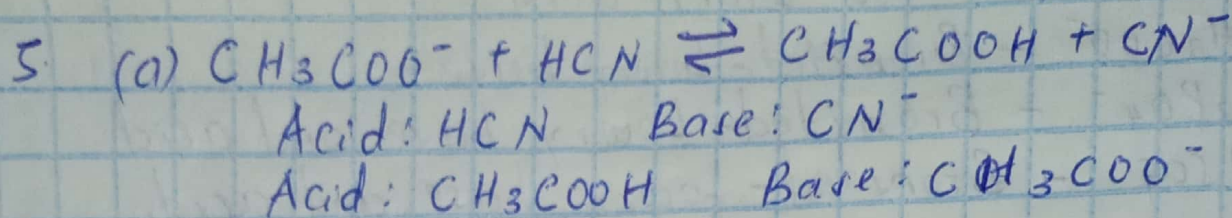
2. An Arrhenius acid is a compound that when dissolved in water will donate an H^+ in solution. An Arrhenius base is a compound that when dissolved in water will yield an OH^- in solution.

A Bronsted-Lowry acid is a molecule that donates a hydrogen ion in a reaction.

A Bronsted-Lowry base is a molecule that accepts hydrogen ion in a reaction.

- 3 (a) HI - Bronsted-Lowry acid
 (b) CH_3COO^- - Bronsted-Lowry base
 (c) H_2PO_4^- - Bronsted-Lowry base and acid
 (d) HSO_4^- - Bronsted-Lowry base and acid.





6. (a) $\text{pH} = -\log [\text{H}^+] = -\log (0.001)$
 $= 3$

(b) $\text{pH} = 14 + \log [\text{H}^+]$
 $= 14 + \log [0.76] = 13.88$

7. (a) $\text{pH} = 2.42$

$\text{pH} = -\log [\text{H}^+]$

$[\text{H}^+] = 10^{-2.42} = 3.801 \times 10^{-3} \text{ mol/L}$

$$(b) \text{ pH} = 11.21$$

$$[\text{H}^+] = 10^{-11.21}$$

$$= 6.166 \times 10^{-12} \text{ mol/L}$$

$$(c) \text{ pH} = 6.96$$

$$[\text{H}^+] = 10^{-6.96}$$

$$= 1.097 \times 10^{-7} \text{ mol/L}$$

$$(d) \text{ pH} = 15$$

$$[\text{H}^+] = 10^{-15}$$

$$= 1.0 \times 10^{-15} \text{ mol/L}$$

$$8. \text{ Volume} = 546 \text{ mL}$$

$$\text{pH} = 10$$

$$\text{pOH} = 14 - \text{pH} = 14 - 10 = 4$$

$$[\text{OH}^-] = 10^{-4} = 1.0 \times 10^{-4} \text{ M}$$

$$\text{moles of NaOH} = \text{Molarity} \times \text{Volume}$$

$$= 1.0 \times 10^{-4} \times 0.546$$

$$= 5.46 \times 10^{-5} \text{ moles}$$

$$\text{Mass of NaOH} = \text{moles of NaOH} \times \text{molar mass}$$

$$= 5.46 \times 10^{-5} \times 40$$

$$= 2.184 \times 10^{-3} \text{ g}$$

9. (a) HNO_3 - strong acid
 (b) HF - weak acid
 (c) H_2SO_4 - strong acid
 (d) HSO_4^- - weak acid
 (e) H_2CO_3 - weak acid
 (f) HCO_3^- - weak acid
 (g) HCl - strong acid
 (h) HCN - weak acid
 (i) HNO_2 - weak acid



(a)	HA	H^+	A^-
initial	0.1	0	0
change	-x	+x	+x
Equilibrium	$0.1 - x$	x	x

The concentration of H^+ would be less than 0.1.
 Therefore, the pH of the solution would be greater than 1.

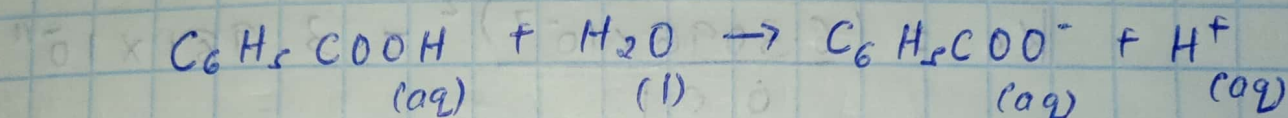
The statement is false

(b) On dissociation, 1 mole of HA gives 1 mole of H^+ and 1 mole of A^- .
 The statement is false

(c) This statement is true since one mole of HX dissociates to give one mole of H^+ and one mole of A^- .

(d) H^+ concentration is less than 0.1. Therefore, the pH is greater than 1. The statement is false

	C_6H_5COOH	$C_6H_5COO^-$	H^+
Initial	0.1	0	0
Change	-x	+x	+x
Equilibrium	0.1-x	x	x



$$K_a = \frac{[C_6H_5COO^-][H^+]}{[C_6H_5COOH]} = \frac{(x)(x)}{0.1-x} = 6.5 \times 10^{-5}$$

Assume $0.1 \gg x$

$$\frac{x^2}{0.1} = 6.5 \times 10^{-5}$$

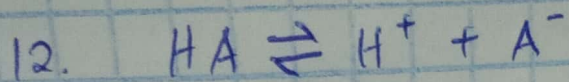
$$x^2 = 6.5 \times 10^{-6}$$

$$x = \sqrt{6.5 \times 10^{-6}} = 2.55 \times 10^{-3} M$$

$$[H^+] = 2.55 \times 10^{-3}$$

$$pH = -\log[H^+] = -\log[2.55 \times 10^{-3}]$$

$$= 2.593$$

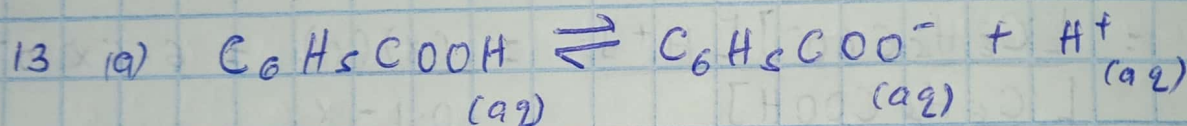


$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$[H^+] = 10^{-pH} = 10^{-6.2} = 6.3096 \times 10^{-7} M$$

$$[H^+] = [A^-]$$

$$K_a = \frac{(6.3096 \times 10^{-7})^2}{0.01} = 3.981 \times 10^{-10}$$



$$K_a = 6.5 \times 10^{-5}$$

Initial	0.2	0	0
Change	-x	+x	+x
Equilibrium	0.2 - x	x	x

$$K_a = \frac{[C_6H_5COO^-][H^+]}{[C_6H_5COOH]}$$

$$\frac{(x)(x)}{0.2 - x} = 6.5 \times 10^{-5}$$

$$x^2 = 6.5 \times 10^{-5} (0.2 - x)$$

$$x^2 = 1.3 \times 10^{-5} - 6.5 \times 10^{-5} x$$

$$x^2 + 6.5 \times 10^{-5}x - 1.3 \times 10^{-5} = 0$$

$$x = 0.003573 \text{ M}$$

$$\% \text{ ionization} = \frac{[H^+]}{0.2} \times 100$$

$$= \frac{0.003573}{0.2} \times 100$$

$$= 1.79\%$$

$$(b) \quad K_a = \frac{[C_6H_5COO^-][H^+]}{[C_6H_5COOH]}$$

$$\frac{(x)(x)}{0.0002 - x} = 6.5 \times 10^{-5}$$

$$x^2 = 6.5 \times 10^{-5} (0.0002 - x)$$

$$x^2 = 1.3 \times 10^{-8} - 6.5 \times 10^{-5}x$$

$$x = 8.606 \times 10^{-5} \text{ M}$$

$$\% \text{ ionization} = \frac{8.606 \times 10^{-5}}{0.0002} \times 100$$

$$= 43.03\%$$

$$14. \quad K_a = \frac{[H^+][A^-]}{[HA]}$$

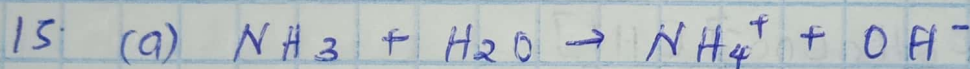
$$\% \text{ ionization} = \frac{[H^+]}{[HA]} \times 100$$

$$\frac{[H^+]}{0.04} = 14\%$$

$$[H^+] = 14(0.04) = 0.0056 \text{ M}$$

$$[H^+] = [A^-]$$

$$K_a = \frac{(0.0056)^2}{0.04 - 0.0056} = 9.12 \times 10^{-4}$$



Initial	0.1	0	0
change	-x	+x	+x
Equilibrium	0.1-x	x	x

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

$$\frac{(x)(x)}{(0.1-x)} = 1.8 \times 10^{-5}$$

$$0.1 \gg x$$

$$\frac{x^2}{0.1} = 1.8 \times 10^{-5}$$

$$x^2 = 1.8 \times 10^{-6}$$

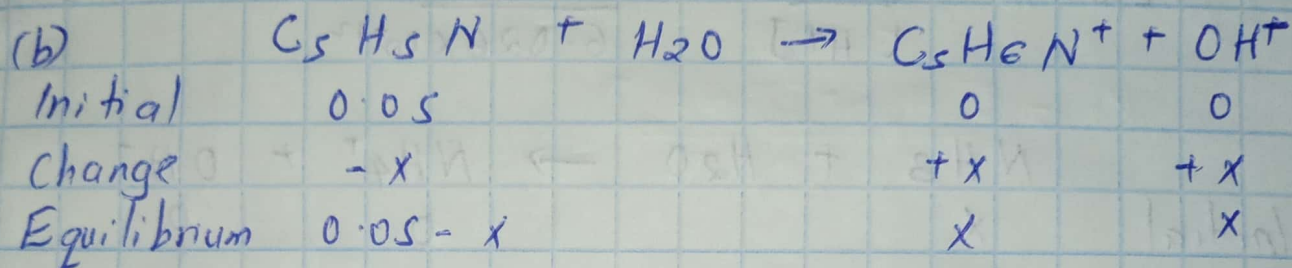
$$x = \sqrt{1.8 \times 10^{-6}} = 1.342 \times 10^{-3} = [OH^-]$$

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

$$= -\log [1.342 \times 10^{-3}]$$

$$= 2.872$$

$$pH = 14 - 2.872 = 11.13$$



$$K_b = \frac{[C_5H_6N^+][OH^-]}{[C_5H_5N]} = 1.7 \times 10^{-9}$$

$$1.7 \times 10^{-9} = \frac{x^2}{0.05 - x}$$

$$0.05 \gg x$$

$$1.7 \times 10^{-9} = \frac{x^2}{0.05}$$

$$x^2 = 8.5 \times 10^{-11}$$

$$x = \sqrt{8.5 \times 10^{-11}} = 9.22 \times 10^{-6}$$

$$pOH = -\log [9.22 \times 10^{-6}] = 5.035$$

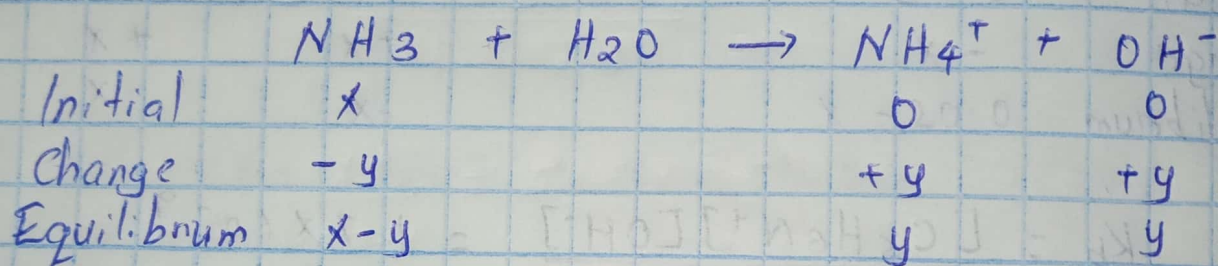
$$pH = 14 - 5.035 = 8.965$$

16. $pH = 11.22$

$$[H^+] = 10^{-11.22} = 6.026 \times 10^{-12} M$$

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

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{6.026 \times 10^{-12}} = 1.66 \times 10^{-3} M$$

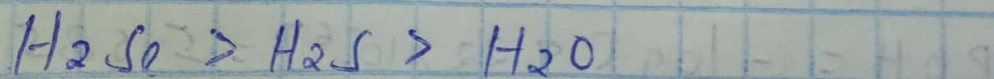


$$K_b = \frac{[y][y]}{[x-y]} = 1.8 \times 10^{-5}$$

$$\frac{(1.66 \times 10^{-3})^2}{x - 1.66 \times 10^{-3}} = 1.8 \times 10^{-5}$$

$$x = 0.155 M$$

17. Electronegativity decreases down the group. Therefore, the bond enthalpy decreases down the group. O has the highest bond enthalpy. Se has the least bond enthalpy.



18 Salt hydrolysis is the reaction between an anion, or cation of a salt with water. Salts are electrolytes that dissociate completely in water.

- (a) Salts that produce basic solutions
- (b) Salts that produce acidic solutions.
- (c) Salts that produce neutral solutions.

19.

	CH_3COONa	H_2O	CH_3COOH	OH^-
Initial	0.36		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	$0.36 - x$		x	x

$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COONa}]}$$

$$\frac{[x][x]}{0.36 - x} = 5.6 \times 10^{-10} = \frac{x^2}{0.36}$$

$$x^2 = 2.016 \times 10^{-10}$$

$$x = 1.42 \times 10^{-5} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log [1.42 \times 10^{-5}] = 4.85$$

$$\text{pH} = 14 - 4.85 = 9.15$$

20. $K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$

	NH_4^+	H_2O	H_3O^+	NH_3
Initial	0.42		0	0
Change	-x		+x	+x
Equilibrium	0.42-x		x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]}$$

$$\frac{x^2}{0.42-x} = 5.6 \times 10^{-10}$$

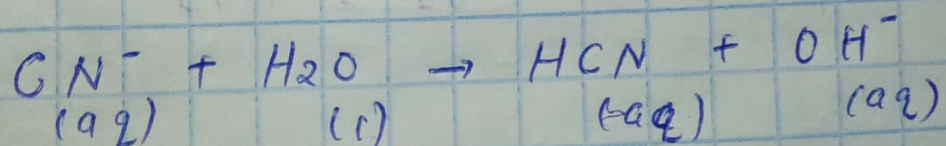
$$x = 1.53 \times 10^{-5} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [1.53 \times 10^{-5}]$$

$$= 4.82$$

21

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.0 \times 10^{-5}$$



$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]}$$

$$\frac{(1.0 \times 10^{-4})^2}{x - 1.0 \times 10^{-4}} = 2 \times 10^{-5}$$

$$x = 6 \times 10^{-4} M = [CN^-]$$

$$\begin{aligned} \text{Moles of } HCN &= M \times V \\ &= 6 \times 10^{-4} \times 0.25 \\ &= 1.5 \times 10^{-4} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Mass} &= \text{Molar mass} \times \text{moles} \\ &= 49.01 \times 1.5 \times 10^{-4} \\ &= 7.35 \times 10^{-3} \text{ g} \end{aligned}$$