If **ΔH°** for the following reaction **is 160.2 kJ/mol**, Use the given data at 298 K to **calculate ΔG°** for the reaction:

$$2\,Cl_2(g) + SO_2(g) \rightarrow SOCl_2(g) + Cl_2O(g)$$

Substance	$Cl_2(g)$	$SO_2(g)$	$SOCl_2(g)$	$Cl_2O(g)$
S° (J/K.mol)	223.0	248.1	309.77	266.1

- a. 195.4 kJ
- b. 186.0 kJ
- c. 135.8kJ
- d. 133.6 kJ
- e. 129.3 kJ

1) $\Delta H^\circ = 160.2 \frac{KJ}{mol}$, $T = 298\ K$

$2Cl_2 + SO_2 \longrightarrow SOCl_2 + Cl_2O$

S° : 223 , 248.1 , 309.77 , 266.1

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$= 160.2 - (298 \times 0.105)$

$= 129 \frac{KJ}{mol.1}$

$\Delta S^\circ = \sum S^\circ_{prod} - \sum S^\circ_{reactant}$

$= 104.77 \frac{J}{K.mol}$

$= 0.104.77 \frac{KJ}{K.mol}$

What is the pH of a 0.35 *M* solution of methylamine (CH_3NH_2, $K_b = 4.4 \times 10^{-4}$) at 25°C?

- a. 1.91
- b. 8.16
- c. 5.55
- d. 12.09
- e. 13.54

2) $[CH_3NH_2] = 0.35\ M$, $K_b = 4.4 \times 10^{-4}$

1. $[OH^-] = \sqrt{K_b\ [base]}$

$= \sqrt{4.4 \times 10^{-4} \times 0.35} = 0.0124\ M$

2. $P_{OH} = -\log [OH^-] = -\log [0.0124] = 1.9$

3. $P_H = 14 - P_{OH} = 14 - 1.9 = 12.1$

What is the **pH** of a 0.03 M solution of NH_3? pK_a for NH_4^+ is 9.26.

- a. 10.9
- b. 10.3
- c. 11.3
- d. 7.0
- e. 12.2

3) $[NH_3] = 0.03$, P_{Ka} for $NH_4^+ = 9.26$

1. $K_a = 10^{-P_{Ka}} = 10^{-9.26} = 5.5 \times 10^{-10}$

2. $K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{5.5 \times 10^{-10}} = 1.82 \times 10^{-5}$

3. $[OH^-] = \sqrt{K_b [base]} = \sqrt{1.82 \times 10^{-5} \times 0.03}$

$= 0.00074$

4. $P_{OH} = -\log [OH^-] = 3.13$

5. $P_H = 14 - P_{OH} = 14 - 3.13 = 10.87$

≈ 10.9

You have prepared 250 mL of an acetate buffer solution having the following concentrations:

$[HC_2H_3O_2] = 0.060$ M, $[C_2H_3O_2^-] = 0.040$ M, What is the **pH** of the buffer? $pK_b = 9.26$ for acetic acid at 25°C.

- a. 4.56
- b. 4.92
- c. 5.69
- d. 9.44
- e. 9.08

4) $[acid] = 0.06$, $[conj] = 0.04$, $pK_b = 9.26$

1. $K_b = 10^{-pK_b} = 10^{-9.26} = 5.5 \times 10^{-10}$

2. $K_a = \frac{K_w}{K_b} = 1.82 \times 10^{-5}$

3. $[H_3O^+] = K_a \frac{[acid]}{[conj]}$

$= 0.0000273$

4. $p_H = -\log [H_3O^+] = 4.56$

At equilibrium, a 1.0 litre container was found to contain 0.20 mol of A, 0.20 mol of B, 0.40 mol of C, and 0.40 mol of D. If 0.10 mol of C and 0.10 mol of D are added to this system, which of the following statements is TRUE?

$$A(g) + B(g) \rightleftharpoons C(g) + D(g)$$

- a. More products are formed to increase Q
- b. More reactants are formed to increase Q
- c. The system remains in equilibrium.
- d. Q=Kc
- e. Q < Kc

The answer is A

The pH of a solution of sodium cyanide (NaCN) in water will be:

- a. Neutral
- b. Less than 7
- c. More than 7
- d. We cannot know

The answer is C

At a certain temperature, the silver-ion concentration in a saturated solution of silver(I) sulfate (Ag_2SO_4) is 2.9×10^{-2} *M*. What is K_{sp} for silver(I) sulfate at this temperature?

- a. 9.6×10^{-5}
- b. 6.9×10^{-7}
- c. 1.2×10^{-5}
- d. 2.1×10^{-4}
- e. 8.3×10^{-4}

$Ag = 2.9 \ 10^{-2} \ M$

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$2Ag + SO_4 \longrightarrow Ag_2SO_4$

2 Mol Ag $\longrightarrow$ 1 Mol SO_4

$\therefore [SO_4] \times 2 = [Ag]$

$$K_{SP} = [Ag]^2 [SO_4] = \frac{[Ag]^3}{2} = 1.2 \ 10^{-5} \ M$$

(c)

Calculate the **solubility** of Zinc Hydroxide ($Zn(OH)_2$, Molar Mass=99.42).

(K_{sp} $Zn(OH)_2 = 4.5 \times 10^{-17}$)

- a. 2.2×10^{-2} g/L
- b. 2.2×10^{-8} g/L
- c. 8.3×10^{-6} g/L
- d. 2.2×10^{-6} g/L
- e. 2.2×10^{-4} g/L

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Q 8) answer is (e)

$$Zn(OH)_2 \longrightarrow Zn + 2OH$$

$$K_{sp} = [Zn][OH]^2$$

$$4.5 \times 10^{-17} = x\,(2x)^2$$

$$4.5 \times 10^{-17} = 4x^3$$

$$x = 2.2 \times 10^{-6} \text{ mol/L}$$

$$\frac{2.2 \times 10^{-6} \text{ mol}}{\text{L}} \;\Big|\; \frac{99.42\text{g}}{1\text{ mol}}$$

$$= 2.2 \times 10^{-4} \text{ g/L}$$

A certain compound has an enthalpy of vaporization equal 52.25 kJ and entropy of vaporization equal 122 J/K at its boiling point. **Calculate the boiling point** (in K) of acetic acid.

- a. 238 K
- b. 0.43 K
- c. 2.3 K
- d. 428 K
- e. 563 K

9) $\Delta S = \frac{\Delta H}{T} \rightarrow T = \frac{\Delta H}{\Delta S} = \frac{52.25\ KJ}{0.122\ KJ} \cdot K$

$= 428\ K$

Which of the following is correct?

- a. A Lewis acid is a species that produces H_3O^+ ion
- b. A Lewis base is a species that produces OH^- ion
- c. A Bronsted-Lowry acid is a species that accepts OH^- ion
- d. A Bronsted-Lowry base is a species that accepts H_3O^+ ion

The answer is D

- e. A Bronsted-Lowry acid is a species that accepts OH^- ion

A sample of 5.40 g of potassium chloride (KCl) is dissolved in 54.60 g of water. What is the **percent by mass** of KCl in the solution?

- a. 22.0%
- b. 16.0%
- c. 8.4%
- d. 9.0%
- e. 92.8%

11) Sample 5.4 g KCl

H_2O 54.6 g

$$\frac{5.4}{5.4 + 54.6} * 100\% = 9\%$$

Which of the following results in a **decrease** in entropy?

- a. $H_2O(s) \rightarrow H_2O(l)$
- b. $N_2(aq) \rightarrow N_2(g)$
- c. $NH_3(l)$, 34.5°C $\rightarrow$ $NH_3(g)$, 34.5°C
- d. Dissolving NaCl in water
- e. $O_2(g)$, 400 K $\rightarrow$ $O_2(g)$, 300 K

The answer is E

Consider the reaction:

$$3Fe(s) + 4H_2O(g) \rightleftharpoons 4H_2(g) + Fe_3O_4(s)$$

If the total pressure is increased suddenly by decreasing the volume, then:

- a. More Fe(s) is produced.
- b. The reaction will proceed from left to right.
- c. The equilibrium constant increases

The answer is D

- d. No change occurs.
- e. More $H_2(g)$ is produced.

If we have the following values for reduction potential

Reduction Half reaction	E°(V)
$Cl_{2(aq)} + 2e^- \rightarrow 2Cl^-_{(aq)}$	1.36
$I_{2(g)} + 2e^- \rightarrow 2I^-_{(aq)}$	0.54
$Br_{2(s)} + 2e^- \rightarrow 2Br^-_{(aq)}$	1.09

Which is the BEST reducing agent?

- a. Chlorine
- b. Iodine
- c. Bromine

The answer is c

- d. Cannot determine from given information

A strip of iron is placed in a 1 *M* solution of iron(II) sulfate, and a strip of copper is placed in a 1 *M* solution of copper(II) chloride. The two solutions are connected with a salt bridge, and the two metals are connected by a wire.

Reduction Half-Reaction	E° (V)
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	−0.41
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	0.34

Which of the following takes place?

- a. Copper atoms will deposit at the cathode
- b. The Fe(II) concentration of the iron half-cell decreases
- c. Copper ions will increase
- d. Iron will deposit at the iron electrode.
- e. No spontaneous reaction will occur

The answer is A

Which one of the following pairs of compounds have a **dipole-dipole** intermolecular forces.

- a. F_2 and C_6H_6
- b. I_2 and NO_3^-
- c. CH_4 and H_2O
- d. NH_3 and CCl_4
- e. HF and H_2O

The answer is E

What is **the boiling point of a solution** that contains 100.0 g of glucose ($C_6H_{12}O_6$; molar mass = 180.0 g/mol) in 900.0 g of H_2O? k_b for water is 0.52 °C/*m*.

- a. 100.79 °C
- b. 100.32 °C
- c. 98.58 °C
- d. 99.68 °C
- e. 100.58 °C

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* Q13) answer is (b)

$m = \frac{n}{m(kg)}$ $\quad n_{C_6H_{12}O_6} = \frac{100}{180} = .56$

$m = \frac{.56}{.9} = .62$ mol / kg

$\Delta T = k_b m$

$100 - T_b = .52 \times .62$

$T_b = 100.322$

For a reaction to be **unfavorable** in all conditions, which of the following statements is correct:

- a. ΔH is positive, ΔG is positive, and ΔS is negative
- b. ΔH is negative, ΔG is negative, and ΔS is positive
- c. ΔH is negative, ΔG is positive, and ΔS is negative
- d. ΔH is negative, ΔG is positive, and ΔS is positive
- e. ΔH is negative, ΔG is negative, and ΔS is negative

The answer is A

Which of the following increases the **solubility** of a gas in a given solvent?

- a. Decreasing the partial pressure of the gas.
- b. Increasing the temperature
- c. Increasing the temperature of the solvent and decreasing the partial pressure of the gas simultaneously
- d. Increasing the partial pressure of the gas.
- e. Adding another gas

The answer is D

For the reaction: $2A + B \rightarrow C + D$ the rate law is:

Rate = $k[A]^2$ and $k = 1.0 \times 10^{-3}$ $L.mol^{-1}.s^{-1}$

If the initial concentration of A is 0.200 M, how long **(in hours)** will it take for the concentration to become 0.0200 M?

- a. 17.1
- b. 25.0
- c. 11.1
- d. 40.0
- e. 12.5

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(Q20) answer is (e)

$$\frac{1}{[A]} = kt + \frac{1}{[A_0]}$$

$$\frac{1}{.02} = 1\times10^{3}\,t + \frac{1}{.2}$$

$$45 = 10^{-3}t \Rightarrow t = 45000\ s$$

$$t = 12.5$$

A galvanic cell consists of a strip of iron that is placed in a 1 *M* solution of iron(II) sulfate and a strip of copper is placed in a 1 *M* solution of copper(II) chloride. Calculate the **standard cell potential** (emf) of this cell at 25°C.

Reduction Half-Reaction	E° (V)
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	−0.41
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	0.34

a. 0.07 V

b. -0.07 V

c. 0.75 V

d. 0.41 V

e. -0.75 V

Anode → Fe^{+2} + 2e → Fe −0.44

cathod → Cu^{+2} + 2e → Cu 0.34

E = 0.34 − (−0.44) = 0.75

The answer is C