

# Chem 209 Final Booklet Solutions

## Solutions to Equilibrium Practice Problems

Problem 3. Solution:

The expression for  $K = \frac{[P_4O_{10}]_{eq}}{[P_4]_{eq}[O_2]_{eq}^5} = 3$

In (a)  $Q = \frac{[P_4O_{10}]_{init}}{[P_4]_{init}[O_2]_{init}^5} = \frac{1 M}{1 M \cdot (1 M)^5} = 1 < 3$ , the reaction proceeds to the right.

In (b)  $Q = \frac{[P_4O_{10}]_{init}}{[P_4]_{init}[O_2]_{init}^5} = \frac{5.0 M}{8.0 M \cdot (0.70 M)^5} = 3.7 > 3$ , the reaction proceeds to the left.

Problem 4. Solution:

To determine the final concentrations, the first thing needed are the initial reactant concentrations and an expression for the reaction coefficient Q.

$$[A_{2(g)}] = \frac{1.00 \text{ mol}}{0.250 \text{ L}} = 4.00 \text{ M}, \quad [AB_{(g)}] = [B_{2(g)}] = \frac{2.00 \text{ mol}}{0.250 \text{ L}} = 8.00 \text{ M},$$

$$\text{and } Q = \frac{[AB]_{init}^2}{[A_2]_{init} [B_2]_{init}} = \frac{(8.00)^2}{(4.00)(8.00)} = 2.0$$

The direction of the reaction needs to be determined. To do this, we compare Q vs K.

Since  $Q = 2.0 > K = 0.5$ , the reaction proceeds to the left.

To determine what final concentrations will be from initial concentrations, a handy tool – the Initial, Change, Equilibrium (ICE) table can be used.

All compounds involved in the reaction are included in an ICE table as follows, with the species that will be consumed on the left side, and the species that will be produced on the right side. Since the reaction proceeds to the left, A<sub>2</sub> and B<sub>2</sub> will be formed and AB will be consumed.

Concentration (M)	2 AB <sub>(g)</sub>	A <sub>2(g)</sub>	B <sub>2(g)</sub>
Initial	8.00 M	4.00 M	8.00 M
Change	-2x	+x	+x
Equilibrium	8.00 - 2x	4.00 + x	8.00 + x

As the reaction proceeds, x moles of A<sub>2(g)</sub> and B<sub>2(g)</sub> are formed as 2x moles of AB<sub>(g)</sub> are consumed. Make sure to consider stoichiometric coefficients appropriately.

The equilibrium values are simply the sum of the initial + change concentrations.

Substitute the equilibrium concentrations into K, and solve for x (remember that K is defined for the reaction in the way that it was initially described):

$$K = \frac{[AB]_{eq}^2}{[A_2]_{eq} [B_2]_{eq}} = \frac{(8.00 - 2x)^2}{(4.00 + x)(8.00 + x)} = 0.5$$

so,

$$(8.00 - 2x)^2 = 0.5(4.00 + x)(8.00 + x)$$

$$64 - 32x + 4x^2 = 0.5(32 + 12x + x^2)$$

$$64 - 32x + 4x^2 = 16 + 6x + 0.5x^2$$

$$3.5x^2 - 38x + 48 = 0$$

To solve for  $x$ , the quadratic formula must be used,

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{38 \pm \sqrt{(-38)^2 - 4(3.5)(48)}}{2(3.5)} = 1.46 \text{ or } 9.39,$$

Since  $2 \times 9.39 = 18.78 > 8.00$  (which would leave the equilibrium concentration of AB at equilibrium negative), then the  $x = 1.46$

So, final concentrations are:

$$[A_2] = 4.00 + 1.46 = 5.46 \text{ M}$$

$$[B_2] = 8.00 + 1.46 = 9.46 \text{ M}$$

$$[AB] = 8.00 - 2(1.46) = 5.08 \text{ M}$$

To check, substitute these concentrations into the Equilibrium constant expression,

$$K = \frac{[AB]_{eq}^2}{[A_2]_{eq}[B_2]_{eq}} = \frac{(5.08)^2}{(5.46)(9.46)} = 0.5, \text{ matches up.}$$

Problem 5. Answer:  $K_c = \frac{[NO]^4}{[H^+]^4[NO_3^-]^4}$

Solution: Pure solids and pure liquids and solvents are not included in the expression. Products go over Reactants, and coefficients in the equation are written as superscripts.

Problem 6. Solution:

$$K_c = [CO_{2(g)}]$$

Remember: Solids and liquids are not included in the expression.

Problem 7. Solution: Equation 2 is equal the double and reverse of equation 1 therefore

$$K_{p2} = K_{p1}^{-2} = \left(\frac{1}{0.157}\right)^2 = 40.6$$

Problem 8. Answer: Equilibrium Constant =  $K_1^{-1/2} = \frac{1}{\sqrt{K_1}}$

Solution: The second equation if reversed (1/K) and halved (K1/2). Combining these two gives 1/K1/2.

Problem 9. Answer: The reaction equation is:



The relationship between  $K_c$  and  $K_p$  is:

$K_p = K_c(RT)^{\Delta n \text{ gas}}$  In this case there are 3 moles of gas in the reactants and 2 moles of gas in the products, so  $\Delta n = -1$

So solving for  $K_p$ :

$$K_p = K_c(RT)^{\Delta n \text{ gas}} = (3.3 \times 10^{91}) \left[ \left( 0.0821 \frac{\text{L atm}}{\text{mol K}} \right) (298 \text{ K}) \right]^{-1} = 1.35 \times 10^{90}$$

Problem 10. Solution:

	$\text{N}_{2(g)} +$	$\text{C}_2\text{H}_2(g) \rightleftharpoons$	$2 \text{HCN}_{(g)}$
i:	1.00	1.00	1.00
c:	+x	+x	-2x
e:	1.00+x	1.00 + x	1.00-2x

$\Rightarrow Q=1 > K$  so then rxn goes to the left

$$\frac{(1.00 - 2x)^2}{(1.00 + x)^2} = K_c \Rightarrow \frac{(1.00 - 2x)}{(1.00 + x)} = \sqrt{K_c}$$

$$x = \frac{1 - \sqrt{K_c}}{2 + \sqrt{K_c}} = 0.488$$

Problem 11. Solution:

	$\text{COCl}_2(g)$	$\rightleftharpoons \text{CO}_{(g)}$	$+ \text{Cl}_{2(g)}$
initial:	0.04	0	0
change:	-x	+x	+x
equil:	0.04-x	x	x

$$K_c = \frac{x^2}{0.04 - x}$$

$$x^2 + K_c x - 0.04 K_c = 0$$

$$x = \frac{-K_c + \sqrt{K_c^2 - (4)(0.04)(K_c)}}{2} = 5.29 \times 10^{-3}$$

Problem 12. Solution:  $K_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = 5.10$

	CO +	H <sub>2</sub> O ⇌	CO <sub>2</sub>	+H <sub>2</sub>
I	0.1M	0.1	0.1	0.1
C	-x	-x	+x	+x
E	0.1-x	0.1-x	0.1+x	0.1+x

$$K_c = \frac{(0.1+x)^2}{(0.1-x)^2} = \frac{0.01+0.2x+x^2}{0.01-0.2x+x^2} = 5.10$$

$$0.01+0.2x+x^2 = 0.051-1.02x+x^2$$

$$0 = 0.041 - 1.22x$$

$$1.22x = 0.041$$

$$x = 0.034M$$

Therefore, at equilibrium  $[H_2] = 0.1 + x = 0.1 + 0.034M = 0.134M$

Problem 13. Solution: Only C is true

An increase in volume will shift the equilibrium to the right thus causing an increase in the total moles of CO at equilibrium.

Problem 14. Solution:

- a) shift to the left
- b) no effect
- c) shift to the right
- d) shift to the right

Problem 15. Solution:

The reaction will shift left forming Ni(CO)<sub>4(g)</sub> to reach equilibrium

Problem 16. Answer: C

$$\text{Solution: } Q = \frac{[NO]^2[Cl_2]}{[NOCl]^2} = \frac{(1.2)^2(0.56)}{(1.3)^2} = 0.51. \quad Q=K, \text{ therefore we are already at equilibrium.}$$

## Solutions to Acids and Bases Practice Problems

Problem 17. Solution: Equal volumes of 0.1 M NaF and 0.1 M HF

Adding an acid and the salt of its conjugate base can form a buffer, this is the case in (b).

Problem 18. Solution: 1 and 3.

A buffer normally consists of a weak acid and its conjugate base in roughly equal amounts. The ratio should be no greater than 0.1 to 10 of weak base to weak acid. In 1, the ratio of acid to base is 3:1. In 3 the ratio is 4:1. In 2, there is a greater amount of strong base than acid, so all of the acetic acid is consumed so it is not a buffer. 4 is not a buffer solution since all of the weak acid is consumed with strong base, 5 is not a buffer because sodium acetate is a weak base, so you have a weak base with a strong base in solution.

Problem 19. Solution: B

This is a buffer solution, therefore:

$$pH = pK_a + \log \left\{ \frac{[A^-]_{eq}}{[HA]_{eq}} \right\} = -\log(1.20 \times 10^{-2}) + \log \left\{ \frac{0.400}{0.600} \right\} = 1.74$$

Problem 20. Solution: This is a buffer solution, therefore:

$$pH = pK_a + \log \left\{ \frac{[A^-]_{eq}}{[HA]_{eq}} \right\} = 4.75 + \log \left\{ \frac{[CH_3COONa]}{0.100} \right\} = 4.75$$

$$\text{Therefore } \log \left\{ \frac{[CH_3COONa]}{0.100} \right\} = 0 \text{ and } \left\{ \frac{[CH_3COONa]}{0.100} \right\} = 1$$

$$[CH_3COONa] = 0.10M$$

$$\# \text{ moles } CH_3COONa = 0.10 \text{ moles}$$

Problem 21. Solution: For this equilibrium  $[H^+] = [In^-] = 10^{-8} = 1 \times 10^{-8}$

$$K_a = 1 \times 10^{-6} = \frac{[H^+][In^-]}{[HIn]} = \frac{(1 \times 10^{-8})^2}{[HIn]}$$

$$[HIn] = 1 \times 10^{-10}, \text{ therefore } [HIn]/[In^-] = 1 \times 10^{-10}/1 \times 10^{-8} = 0.01 = 1/100$$



Problem 22. Solution: pH = 8.91

$$\begin{aligned} \text{pH} &= \text{pKa} + \log \frac{\text{CN}^-}{\text{HCN}} \\ &= 9.20 + \log \frac{1.0}{2.0} \\ &= 8.91 \end{aligned}$$

Problem 23. Solution: C

This is a buffer system as there are equal amounts of conjugate acid and base, but the addition of H<sup>+</sup> can change the pH slightly. The original pH of the buffer is approximately equal to pKa. When H<sup>+</sup> ions are added from the strong acid HCl, A<sup>-</sup> is converted into HA. Therefore the addition of 0.01 moles H<sup>+</sup> produces 0.01 moles HA and consumes 0.01 moles A<sup>-</sup>.

$$n\text{HA} = 0.5 + 0.01 = 0.51 \text{ mol}$$

$$n\text{A}^- = 0.5 - 0.01 = 0.49 \text{ mol}$$

$$[\text{H}^+] = K_a \times n\text{HA}/n\text{A}^- = (1.8 \times 10^{-4}) (0.51/0.49) = 1.873 \times 10^{-4}$$

$$\text{pH} = -\log (1.873 \times 10^{-4}) = 3.727$$

Problem 24. Solution: C

There are more moles of carbonate buffer in solution c than any of the other solutions

Problem 25. Solution: E

The solution is a buffer with pH above 7. A buffer is resistant to both addition of strong acid and strong base and the concentration of the hydronium ion is not more than the hydroxide ion (pH > 7).

Problem 26. Solution: E

	HCN +	H <sub>2</sub> O ⇌	CN <sup>-</sup> +	H <sub>3</sub> O <sup>+</sup>
I	0.5		y	0
C	-1 × 10 <sup>-7</sup>		+1 × 10 <sup>-7</sup>	+1 × 10 <sup>-7</sup>
E	0.5		y + 1 × 10 <sup>-7</sup>	+1 × 10 <sup>-7</sup>

$$K_a = 6.2 \times 10^{-10} = [(1 \times 10^{-7})(Y + 1 \times 10^{-7})] / 0.5$$

$$Y = 0.003\text{M} \times 1\text{L} = 0.003\text{moles}$$

$$\text{Mass} = 49.0075\text{g/mol NaCN} \times 0.003 \text{ moles} = 0.15\text{g}$$



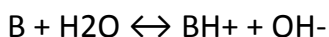
Problem 27. Solution: E

A buffer normally consists of a weak acid and its conjugate base in roughly equal amounts. The ratio should be no greater than 0.1 to 10 of weak base to weak acid. In 1, the ratio of acid to base is 3:1. In 3 the ratio is 4:1. In 2, there is a greater amount of strong base than acid.

Problem 28. Solution: C or E

In order to determine which solutions are able to act as buffer solutions, determine what ions will be found in the solution and whether those ions are acidic, basic or spectator ions. If acidic and/or basic ions are found, calculate the amount of ions that are present. For solution A) all ions present ( $\text{H}^+$ ,  $\text{NO}_3^-$ , and  $\text{Na}^+$ ) are spectator ions. No buffer abilities possible. For solution B) all ions present ( $\text{Na}^+$ ,  $\text{OH}^-$  and  $\text{Cl}^-$ ) are spectator ions. No buffer abilities possible. For solution C) This is a 1:1 ratio of conjugate acid to its conjugate base. This IS a buffer. For solution D) the  $\text{HCl}$  completely dissociates to form  $\text{H}^+$  and  $\text{Cl}^-$  ions.  $\text{NH}_3$  can form an equilibrium where  $\text{NH}_3 + \text{H}^+ \leftrightarrow \text{NH}_4^+$ . However,  $\text{HCl}$  and its dissociated ions are present in 0.01mol amounts—the same as the amount of  $\text{NH}_3$ . Therefore, all the  $\text{NH}_3$  is “used up” in reacting with the  $\text{HCl}$ , so no buffering abilities are possible. For solution E) the  $\text{NaOH}$  completely dissociates to form  $\text{Na}^+$  and  $\text{OH}^-$  ions. The 0.004mol of  $\text{OH}^-$  will react with the 0.01mol of  $\text{HF}$  to form 0.004mol of  $\text{F}^-$  and have 0.006mol left of  $\text{HF}$  since the  $\text{OH}^-$  is the limiting reagent. This means we have  $\text{HA}$  and  $\text{A}^-$  both present in a ratio of 3:2. This is a buffer.

Problem 29. Solution: C



Calculate from the given pH, the concentration of  $\text{OH}^-$  ions that dissociate form from the reaction of the base with water.

$$\text{pOH} = 14 - \text{pH} = 14 - (8.88) = 5.12$$

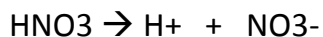
$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-5.12} = 7.585 \times 10^{-6}\text{M}$$

Assume 1 L of solution, therefore the  $[\text{B}] = 0.40\text{mol/L}$  and  $[\text{BH}^+] = 0.250\text{mol/L}$ .

$$K = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]} = \frac{(0.250)(7.585 \times 10^{-6})}{0.40} = 4.74 \times 10^{-6}$$

Problem 30. Solution: D

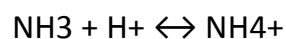
Because equal volumes of the acid and weak base are being mixed, all concentrations (M) can be treated as moles (mol). HNO<sub>3</sub> is a strong acid so it completely dissociates



0.1      0.1      0.1

NH<sub>3</sub> will react with the H<sup>+</sup> released by the HNO<sub>3</sub> to form NH<sub>4</sub><sup>+</sup>. Initially there is 0.3M or 0.3moles of NH<sub>3</sub>. Upon addition of 0.1moles H<sup>+</sup> (from the HNO<sub>3</sub>), 0.1mol of NH<sub>3</sub> will react to form 0.1mol of NH<sub>4</sub><sup>+</sup>, leaving 0.2mol NH<sub>3</sub> unreacted.

Therefore,



Using the following equation, solve for [H<sup>+</sup>].

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

[A-]

$$K_w = K_a \times K_b \text{ so that } K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$K_b = 1.8 \times 10^{-5}$$

$$[\text{H}^+] = 5.56 \times 10^{-10} \times \frac{[0.2]}{[0.1]} = 2.78 \times 10^{-10} \text{M}$$

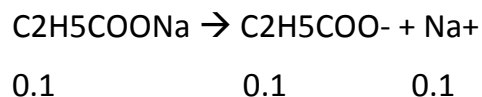
[0.2]

$$\text{pH} = -\log[\text{H}^+] = -\log[2.78 \times 10^{-10}] = 9.56$$

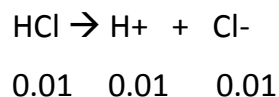
Problem 31. Solution: B

Write out the equations that are occurring in the solution described above. Because both solutes are being added to 1L of water, all concentrations (M) can be treated as moles (mol).

C<sub>2</sub>H<sub>5</sub>COONa is a soluble salt so it completely dissociates



HCl is a strong acid so it completely dissociates



C<sub>2</sub>H<sub>5</sub>COO<sup>-</sup> will react with all the H<sup>+</sup> to form C<sub>2</sub>H<sub>5</sub>COOH. Initially there is 0.1 mole of C<sub>2</sub>H<sub>5</sub>COO<sup>-</sup>. When 0.01mol of H<sup>+</sup> is added, the C<sub>2</sub>H<sub>5</sub>COO<sup>-</sup> reacts leaving 0.09mol. There is also 0.1mol of C<sub>2</sub>H<sub>5</sub>COOH to start, but when the C<sub>2</sub>H<sub>5</sub>COO<sup>-</sup> reacts with the H<sup>+</sup>, it forms 0.01mol more C<sub>2</sub>H<sub>5</sub>COOH so that the total amount of C<sub>2</sub>H<sub>5</sub>COOH is 0.11mol.



Before HCl is added    0.1                      0.1

After HCl is added    0.11                      0.09

Calculate [H<sup>+</sup>] using the following equation. (n = moles)

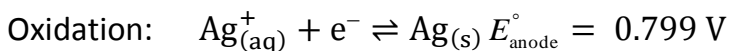
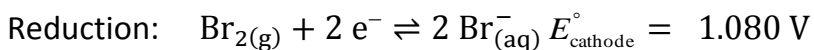
$$[\text{H}^+] = K_a \times \frac{n_{\text{HA}}}{n_{\text{A}^-}} = 1.41 \times 10^{-5} \times \frac{0.11}{0.09} = 1.72 \times 10^{-5} \text{M}$$

Calculate pH from the [H<sup>+</sup>].

$$\text{pH} = -\log[\text{H}^+] = -\log[1.72 \times 10^{-5}] = 4.76$$

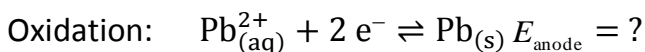
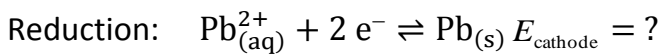
## Solutions to Electrochemistry Practice Problems

Problem 33. Solution: List the oxidation and reduction steps:



$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ $= 1.080 \text{ V} - 0.799 \text{ V} = 0.281 \text{ V}$ $E^\circ_{\text{cell}} = \frac{0.0257 \text{ V}}{n} \ln K_{eq} \quad \text{or} \quad E^\circ_{\text{cell}} = \frac{0.0592 \text{ V}}{n} \log K_{eq}$	$\log K_{eq} = \frac{nE^\circ_{\text{cell}}}{0.0592 \text{ V}}$ $\log K_{eq} = \frac{2(0.281 \text{ V})}{0.0592 \text{ V}} = 9.49$ $K_{eq} = 3.1 \times 10^{+9}$
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Problem 34. Solution: The half-cell reactions are as follows (note: these are not at standard state or  $E_{\text{cell}} = 0 \text{ V}$ ):



$$E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$$

$$E_{\text{cathode}} = -0.125 - \frac{0.0592 \text{ V}}{2} \log \frac{1}{0.100 \text{ M}} = -0.1546 \text{ V}$$

$$E_{\text{anode}} = -0.125 - \frac{0.0592 \text{ V}}{2} \log \frac{1}{[\text{Pb}^{2+}]}$$

$$E_{\text{cell}} = 0.0700 \text{ V} = -0.1546 - \left( -0.125 - \frac{0.0592 \text{ V}}{2} \log \frac{1}{[\text{Pb}^{2+}]} \right)$$

$$[\text{Pb}^{2+}_{(aq)}] = 4.3 \times 10^{-4}$$

Solubility	$\text{PbSO}_4$	$\text{Pb}^{2+}_{(aq)}$	$\text{SO}_4^{2-}$
Initial	Some	0	0
Change	-s	+s	+s
Equilibrium	Some-s	S	S

$$K_{\text{sp}} = [\text{Pb}^{2+}_{(aq)}][\text{SO}_4^{2-}]$$

$$K_{\text{sp}} = (4.3 \times 10^{-4})^2 = 1.9 \times 10^{-7}$$



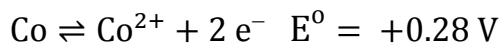
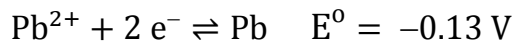
Problem 35. Answers:

Cr = +3, O = -2

Ca = +2, C = +4, H = +1, O = -2

Fe = +3, C = +4, O = -2

Problem 36. a) Answer: 0.15 V



$$E_{\text{cell}}^{\circ} = 0.28 \text{ V} - 0.13 \text{ V} = 0.15 \text{ V}$$

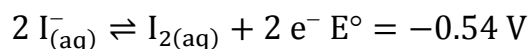
(b) Answer: B

Reduction occurs at the cathode; therefore, the lead electrode is the cathode.

Problem 37. Answer:  $\text{MnO}_4^{-}(\text{aq}) > \text{Zn}(\text{s}) > \text{I}_2(\text{aq}) > \text{I}^{-}(\text{aq}) > \text{Zn}^{2+}(\text{aq}) > \text{MnO}_2(\text{aq})$

An oxidizing agent gets reduced therefore  $\text{MnO}_4^{-}(\text{aq})$  is the strongest oxidizing agent as it has the largest  $E^{\circ}$ .

Problem 38. Solution:

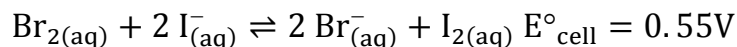
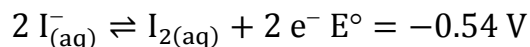
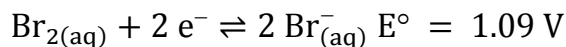


$\text{Br}_2(\text{aq})$  will be reduced and  $\text{I}^{-}(\text{aq})$  will be oxidized.



Problem 39. What is the value of  $E_{\text{cell}}$ ?

Answer: 0.55 V

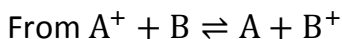


Then put the  $E^\circ_{\text{cell}}$  into the Nerst:

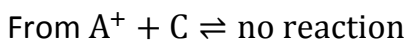
$$E^\circ_{\text{cell}} = 0.55 - \left( \frac{0.0592}{2} \right) \left( \log \left( \frac{[0.2]^2 [0.1]}{[0.2][10]^2} \right) \right) = 0.66 \text{ V}$$

Problem 40. Answer: D>B>A>C

Solution: Reducing agent undergoes oxidation (most easily oxidized = strongest reducing agent)



B can oxidize to  $\text{B}^+$ ,  $\therefore$  B is a stronger reducing agent than A.



C cannot oxidize to  $\text{C}^+$   $\therefore$  A is a stronger reducing agent than C.



D can oxidize to  $\text{D}^{2+}$   $\therefore$  D is a stronger reducing agent than B.

The decreasing order of reactivity (most easily oxidized to least easily oxidized) is:

D>B>A>C, where D is the strongest reducing agent.

Problem 41. Answer: B

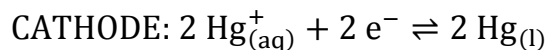
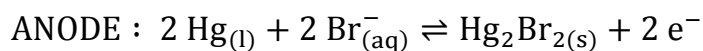


To Find: anode reaction

ANODE allows oxidation to take place, which is the loss of electrons.

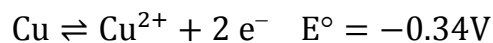
- Answers A, C and E are incorrect as they are reduction reactions.

- Answer D is incorrect as the charges on both sides are NOT balanced.



$2 \text{Hg}_{(\text{l})}$  acts as an intermediate

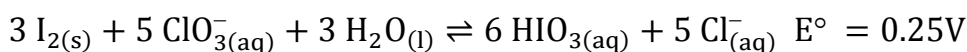
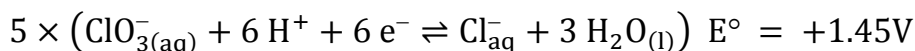
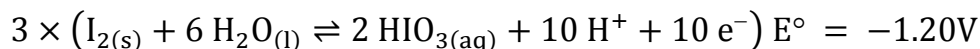
Problem 42. Answer:  $\text{Ag}^{+}$  only



So addition of this to the cathode reaction should be a positive number for the reaction to proceed.

The only one that this works for is  $\text{Ag}^{+}$

Problem 43. Answer: 0.25 V



Problem 44. Answer: 0.62 V

In the reaction above,  $\text{Sn}_{(\text{aq})}^{2+}$  is being oxidized to  $\text{Sn}_{(\text{aq})}^{4+}$  (it lost  $2 \text{e}^{-}$ ) and  $\text{Fe}_{(\text{aq})}^{3+}$  is being reduced to  $\text{Fe}_{(\text{aq})}^{2+}$  (gaining  $1 \text{e}^{-}$ ).  $E_{\text{cell}} = E_{\text{reduction}} + E_{\text{oxidation}}$ . Since all E values are given as  $E_{\text{reduction}}$  values, the  $E_{\text{red}}$  value for Fe in the reaction is  $+0.77 \text{V}$ . The E value for Sn in the reaction must be reversed because Sn is undergoing oxidation, therefore  $E_{\text{ox}} = -0.15\text{V}$ .

$$E_{\text{cell}} = E_{\text{reduction}} + E_{\text{oxidation}} = (+0.77\text{V}) + (-0.15\text{V}) = 0.62\text{V}$$

Problem 45. Answer: A only

Calculate the  $E_{\text{cell}}$  for all three reactions. In statement I),  $E_{\text{cell}} = -1.86\text{V}$  ( $\text{Cd}_{(\text{aq})}^{2+}$  is being reduced and  $\text{Cl}_{2(\text{g})}$  is being oxidized). In statement II),  $E_{\text{cell}} = -0.29\text{V}$  ( $\text{Sn}_{(\text{aq})}^{2+}$  is being oxidized and reduced into  $\text{Sn}_{(\text{aq})}^{4+}$  and  $\text{Sn}$  respectively). In statement III)  $E_{\text{cell}} = 0.91\text{V}$  ( $\text{Sn}$  is being oxidized and  $\text{Fe}^{3+}$  is being reduced).  $E_{\text{cell}}$  values that are positive means that the redox reactions will occur spontaneously, while a negative  $E_{\text{cell}}$  value means the redox reaction is not spontaneous.

Problem 46. Answer: I, II, and III

Solution: Statement I) is true because  $\text{Cu}^{2+}$  has a more positive  $E_{\text{reduction}}$  value than  $\text{Cr}^{3+}$ , which means that  $\text{Cu}^{2+}$  is a better oxidizing agent than  $\text{Cr}^{3+}$  (remember that an oxidizing agent oxidizes other substances and becomes reduced in the process).

Problem 47. Answer: -0.26

In the reaction given above,  $\text{Ag}_{(\text{aq})}^{+}$  is being reduced to  $\text{Ag}$  while  $\text{Ni}$  is being oxidized to  $\text{Ni}_{(\text{aq})}^{2+}$ . The overall  $E_{\text{cell}}^{\circ}$  for the reaction is calculated as

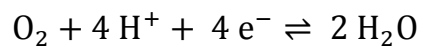
$$E_{\text{cell}}^{\circ} = E_{\text{reduction}} + E_{\text{oxidation}}$$

The reduction half of the reaction is given as  $+0.80\text{V}$ . In order to calculate the  $E$  value for the oxidation half, rearrange the  $E_{\text{cell}}^{\circ}$  equation to solve for  $E_{\text{oxidation}}$ .

$E_{\text{oxidation}} = E_{\text{cell}}^{\circ} - E_{\text{reduction}} = (+1.06\text{V}) - (+0.80\text{V}) = +0.26\text{V}$ . The  $E_{\text{oxidation}}$  value is the opposite sign from the  $E_{\text{reduction}}$  value, so in order to solve for the  $E_{\text{reduction}}$  value as asked, the sign on the  $E_{\text{oxidation}}$  value must be reversed. Therefore,  $E_{\text{reduction}}$  for  $\text{Ni}_{(\text{aq})}^{2+} = -0.26\text{V}$

Problem 48. Answer:  $n = 4$

$2 \text{H}_2 + \text{O}_2 \rightleftharpoons 2 \text{H}_2\text{O}$  can be broken down into its half-cell reactions:



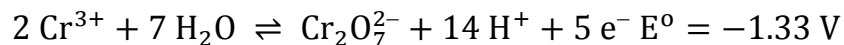
Since there is an exchange of four electrons,  $n = 4$  in the Nernst Equation.

Problem 49. Answer: 0.21 V

$$\text{Solution: } E = E^\circ - \frac{0.0257\text{V}}{n} \ln Q = 0.15\text{V} - \frac{0.0257\text{V}}{2} \ln \frac{[\text{Co}^{2+}]}{[\text{Pb}^{2+}]} = 0.21\text{V}$$

Problem 50. (a) When current is allowed to flow, which species is oxidized?

Solution:



Oxidation is a loss of electrons.  $\text{Cr}^{3+}$  is being oxidized to  $\text{Cr}_2\text{O}_7^{2-}$

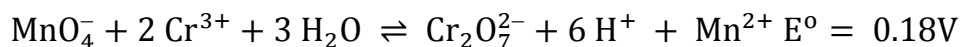
(b) When current is allowed to flow, which species is reduced?

Solution:

$\text{MnO}_4^-$  is being reduced to  $\text{Mn}^{2+}$

(c) What is the value of  $E^\circ_{\text{cell}}$ ?

Solution:



(d) What is the oxidation state of Cr in  $\text{Cr}_2\text{O}_7^{2-}$ ?

Solution: oxid. # =  $[7(-2) + 2]/2 = +6$

## Solutions to Kinetics Practice Problems

Problem 51. Solution:

$$\frac{v_1}{v_2} = \frac{k[I^-]_1^m [S_2O_8^{2-}]_1^n}{k[I^-]_2^m [S_2O_8^{2-}]_2^n}$$

$$\frac{1.25 \times 10^{-5}}{6.25 \times 10^{-6}} = \frac{k(0.080)^m (0.040)^n}{k(0.040)^m (0.040)^n} = 2 = 2^m \Rightarrow m = 1$$

$$\frac{v_1}{v_3} = \frac{1.25 \times 10^{-5}}{6.25 \times 10^{-6}} = \frac{k(0.080)^m (0.040)^n}{k(0.080)^m (0.020)^n} = 2 = 2^n \Rightarrow n = 1$$

$$\text{rate} = k[I^-]^1 [S_2O_8^{2-}]^1$$

$$k = \frac{\text{rate}}{[I^-][S_2O_8^{2-}]}$$

$$k_1 = \frac{1.25 \times 10^{-5}}{(0.080) \cdot (0.040)} = 3.91 \times 10^{-3} \text{ M}^{-1} \cdot \text{s}^{-1}$$

Problem 52. Solution:

$$\text{Rate} = k[\text{H}_2\text{O}]^x [\text{CH}_3\text{Cl}]^y$$

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \frac{k(0.0100)^x}{k(0.0200)^x} = \frac{3.6 \times 10^{-4}}{1.44 \times 10^{-4}}$$

$0.5^x = .25$ , therefore  $x = 2$  and the reaction is second-order in  $\text{H}_2\text{O}$

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \frac{k(0.0100)^x}{k(0.0200)^x} = \frac{3.6 \times 10^{-4}}{1.44 \times 10^{-4}}$$

$$5.54 \left(\frac{2}{3}\right)^y = 3.69$$

$\left(\frac{2}{3}\right)^y = \frac{2}{3}$ , therefore  $y = 1$  and the reaction is first-order in  $\text{CH}_3\text{Cl}$

$$\text{Rate} = k[\text{H}_2\text{O}]^2 [\text{CH}_3\text{Cl}]^1$$

Problem 53. Solution

$$\ln[A]_{t_{1/2}} = \ln[A]_0 - kt_{1/2}, \quad \ln\left(\frac{[A]_0}{[A]_{t_{1/2}}}\right) = kt_{1/2}$$

$$\text{for } t_{1/2} \quad [A]_{t_{1/2}} = \frac{[A]_0}{2}$$

$$\ln\left(\frac{[A]_0}{[A]_{t_{1/2}}}\right) = \ln\left(\frac{[A]_0}{[A]_0/2}\right) = \ln 2 = kt_{1/2}$$

$$\Rightarrow t_{1/2} = \frac{\ln(2)}{k}$$

Problem 54. Solution: If 20% decomposes, then 80% of the sample remains.

$$\ln[A] = \ln[A]_0 - kt \quad \text{rearrange}$$

$$\ln\left(\frac{0.8[A]_0}{[A]_0}\right) = -k(50s)$$

$$k = 4.463 \times 10^{-3} \text{ sec}^{-1}$$

$$t_{1/2} = \frac{\ln 2}{k} = 155 \text{ sec}$$

Problem 55. Solution: (a)

$$\frac{v_1}{v_2} = \frac{k[CH_3NNCH_3]_1^n}{k[CH_3NNCH_3]_2^n}$$

$$\frac{2.8 \times 10^{-6}}{1.1 \times 10^{-5}} = \frac{(5.13 \times 10^{-2})^n}{(2.05 \times 10^{-1})^n} = 0.25 = 0.25^n \Rightarrow n = 1$$

Therefore, the reaction is the first order.

$$\text{Rate} = k [CH_3NNCH_3(g)], \text{ therefore } k = \frac{2.8 \times 10^{-6} \text{ M/s}}{5.13 \times 10^{-2} \text{ M}} = 5.46 \times 10^{-5} \text{ s}^{-1}$$

$$(b) \quad \ln[A]_t = \ln[A]_0 - kt$$

$$\ln[0.1 - 5.13 \times 10^{-2}] = \ln[5.13 \times 10^{-2}] - (5.46 \times 10^{-5}) \times t$$

$$-5.27 = -2.97 - (5.46 \times 10^{-5}) \cdot t$$

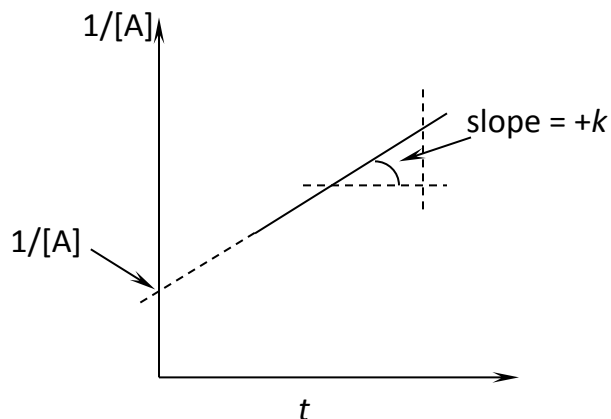
$$(5.46 \times 10^{-5}) \cdot t = 2.30$$

$$t = 4.21 \times 10^4 \text{ s} = 11.7 \text{ hours}$$

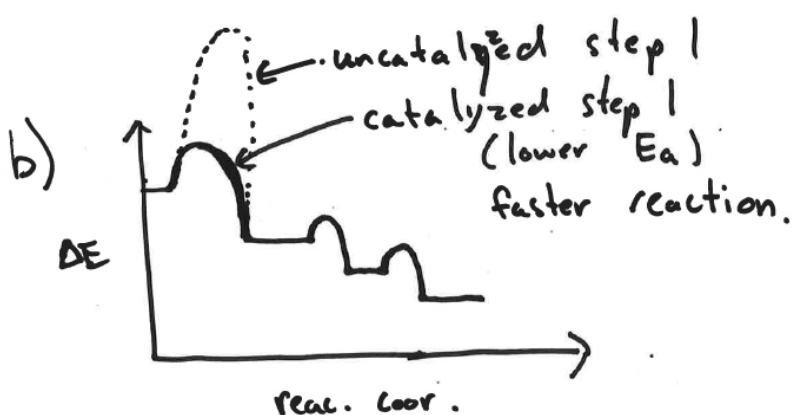
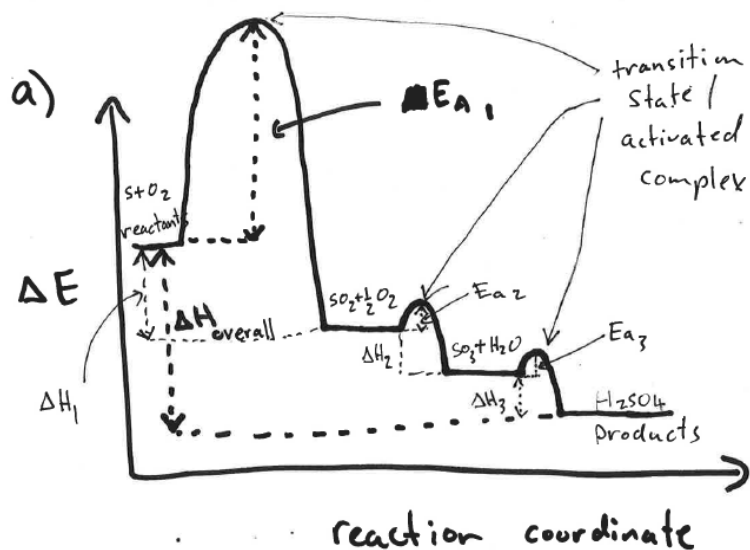
$$(c) \quad [A]_{10\%} = 0.1 \cdot 5.13 \times 10^{-2} = 5.12 \times 10^{-3} \text{ M}$$

$$\text{Rate} = k[A] = (5.46 \times 10^{-5} \text{ s}^{-1})(5.12 \times 10^{-3} \text{ M}) = 2.80 \times 10^{-7} \text{ M/s}$$

Problem 56. Solution: Because this reaction is second order  $\frac{1}{[A]} - \frac{1}{[A]_0} = kt$  if  $1/[A]$  was plotted vs.  $t$  then the y-intercept would be  $1/[A]_0$  and the slope would be  $k$



Problem 57. Answer



Problem 58. Solution: Since  $\Delta H_{rxn} > 0$ , then  $\Delta H_{products} > \Delta H_{reactants}$



Thus,  $E_a \text{ rev} = 66 - 41 = 25 \text{ kJ/mol}$

(draw an energy diagram to see this better)

Problem 59. Solution:

Let  $T_1 = 96^\circ\text{C} = 369 \text{ K}$ , and  $T_2 = 25^\circ\text{C} = 298 \text{ K}$ .

Using the Arrhenius Equation  $\ln\left(\frac{k_1}{k_2}\right) = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$

$$\ln\left(\frac{3.55 \times 10^{10} \text{ s}^{-1}}{1.2 \times 10^{10} \text{ s}^{-1}}\right) = -\frac{E_a}{8.314 \text{ JK}^{-1} \text{ mol}^{-1}} \left(\frac{1}{369 \text{ K}} - \frac{1}{298 \text{ K}}\right)$$

and solve for  $E_a$ ,  $E_a = 12.8 \text{ kJ/mol}$

Problem 60. Solution: from Arrhenius equation  $\ln\left(\frac{k_2}{k_1}\right) = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$

since  $v \propto k$ , then

$$\ln\left(\frac{k_2}{k_1}\right) = \ln\left(\frac{v_2}{v_1}\right) = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$\ln(3) = -\frac{E_a}{8.314} \left(\frac{1}{325} - \frac{1}{313}\right) \Rightarrow E_a = 77433.3 \text{ J/mol} = 77.4 \text{ kJ/mol}$$

Problem 61. Solution: Since we have an elementary process, then rate of the reaction is  $\text{rate} = k[A]^2$ ,

This reaction is second order therefore  $\frac{1}{[A]} - \frac{1}{[A]_0} = kt$

$$\frac{1}{0.1 \text{ M}} - \frac{1}{0.2 \text{ M}} = k(35.2 \text{ min})$$

$$5 \text{ M}^{-1} = k \cdot (35.2 \text{ min})$$

$$k = 1.42 \times 10^{-1} \text{ M}^{-1} \cdot \text{min}^{-1}$$

Problem 62. The steps are elementary, and the first step is the rate determining step. So the overall rate law only depends on the first step:

$$\text{rate} = k[\text{O}_3][\text{NO}]$$

## Periodic Properties Practice Problems

Problem 63. Answer: a)  $V^{5+} < Ti^{4+} < Sr^{2+} < Br^-$

The electron configurations are...

	# of protons
$Sr^{2+}$ : [Kr]	38
$Br^-$ : [Kr]	35
$V^{5+}$ : [Ar]	23
$Ti^{4+}$ : [Ar]	22

$Sr^{2+}$  and  $Br^-$  has the same number of electrons; however,  $Z_{eff}$  is greater for  $Sr^{2+}$  due to a greater # of protons, resulting in a greater + charge that pulls the electrons closer to the nucleus. Thus,  $Sr^{2+}$  is smaller than  $Br^-$ .

$V^{5+}$  and  $Ti^{4+}$  has the same number of electrons; however,  $Z_{eff}$  is greater for  $V^{5+}$  due to a greater # of protons. Thus  $V^{5+}$  is smaller than  $Ti^{4+}$ .

$Sr^{2+}$  and  $Br^-$  are larger than  $V^{5+}$  and  $Ti^{4+}$  because there are more electrons held in a larger subshell. (smallest)  $V^{5+} < Ti^{4+} < Sr^{2+} < Br^-$  (largest)

b)  $Cl > Br > I$

Problem 64. Solutions

a) Electron affinity becomes more negative from left to right because Cl has higher  $Z_{eff}$  than S.

b) For k, valence electron is in 4s orbital while valence electron in Li is in 2s orbital. 4s orbital is larger than 2s.

c) Ionization of Xe removes 5p electron while ionization of Kr removes 4p electron. Ionization energy  $\propto \frac{1}{n^2}$

d)  $Z_{eff}$  is larger in O than in B.

Problem 65.

a) Mg, Ionization energy increase as you move up a group and it also increases as you move right across a period.

b) Mg, Radius increase as you move left along a period and as you move down a group.

c)  $\text{Ca} < \text{Be} < \text{P} < \text{Cl} < \text{O}$

Electronegativity increases as you up a group and as you move right across a period. F is the most electronegative element and Cs is the least electronegative element.

Problem 66. Answer:

- a)  $[\text{Ar}]$  or  $[\text{Ne}]3s^23p^6$
- b)  $\text{S}^{2-}$
- c)  $\text{S}^{2-}$
- d)  $\text{Ca}^{2+}$
- e)  $\text{S}^{2-} < \text{Ar} < \text{Ca}^{2+}$

Solution:

- a)  $[\text{Ar}]$  or  $[\text{Ne}]3s^23p^6$
- b) Ar,  $\text{Ca}^{2+}$  and  $\text{S}^{2-}$  all have the same number of electrons; however, Ar has 18 protons,  $\text{Ca}^{2+}$  has 20, and  $\text{S}^{2-}$  has 16. Thus  $\text{S}^{2-}$  has the least  $Z_{\text{eff}}$  since it has the smallest charge pulling on the electrons.
- c) The species with the least favorable electron affinity is  $\text{S}^{2-}$  because it has the smallest  $Z_{\text{eff}}$  (a smaller positive charge to pull electrons towards the nucleus).
- d)  $\text{Ca}^{2+}$  has the greatest  $Z_{\text{eff}}$  because it has the greatest number of protons pulling on the 18 electrons, resulting in the electrons being closer to the nucleus.

Ionization energy is the energy required to remove an electron. It is most difficult to remove an electron from  $\text{Ca}^{2+}$ , as this will involve the removal of a core electron instead of a valence electron, and  $\text{Ca}^{2+}$  has the greatest  $Z_{\text{eff}}$ . It is easier to remove an electron from  $\text{S}^{2-}$  versus Ar because  $\text{S}^{2-}$  has a smaller  $Z_{\text{eff}}$ .  $\text{S}^{2-} < \text{Ar} < \text{Ca}^{2+}$

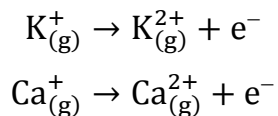
## Problem 67.

has a noble gas electron configuration?	Kr
has the smallest ionization energy?	Cs
has an atomic number $Z = 13$ ?	Al
has a half-filled sub-shell with $l = 2$ ?	Cr
is a hydrogen-like species?	$\text{He}^+$
has only one 4s electron?	Cr
have two unpaired electrons?	O, C, $\text{V}^{3+}$
has only two d electrons with $n = 3$ ?	$\text{V}^{3+}$
is diamagnetic?	Kr
has the largest radius?	Cs
has only one electron with $l = 1$ ?	Al
has the largest number of unpaired electrons?	Cr
are transition metal species?	$\text{V}^{3+}$ , Cr

Problem 68. Answer:

a)  $IE_1$  for K  $>$   $IE_1$  for Ca because  $Z_{\text{eff}}$  increases from left to right across a period, so Ca has a higher  $Z_{\text{eff}}$  and it is therefore harder to remove an electron.  $IE_2$  for K  $>$   $IE_2$  for Ca because the ionization process for Ca leads to the formation of the stable noble gas configuration ( $\text{Ca}^{2+}$ ) while the ionization reaction for K requires the destruction of a stable noble gas configuration. The former will always be lower in energy.

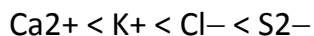
b)



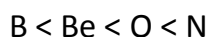
c) The larger the negative charge, the greater the repulsion between electrons, the larger the radius. These are all atoms that have the configuration of Kr, so only charge influences radius. Therefore,  $\text{Rb}^+ < \text{Br}^- < \text{Se}^{2-}$

The ionization energy decreases as you go down Group 1 because as  $n$  increases it is easier to remove the outer electron. Radius increases as you go down Group A because as  $n$  increases, the radius increases.

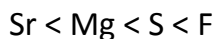
Problem 69. a) Answer:



b) Answer:



c) Answer:



Problem 70. Answer:

a) When AB is put into aqueous solution, it will dissociate into A<sup>-</sup> and B<sup>+</sup> ions rather than A<sup>+</sup> and B<sup>-</sup>. A has a greater ionization energy, and thus is less able to form A<sup>+</sup> than B forming B<sup>+</sup>. Further, electron affinity for A is larger, and thus the reaction A → A<sup>-</sup> is more favorable than B → B<sup>-</sup>.

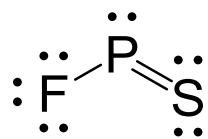
b) A has a greater electronegativity; A has a larger electron affinity value and thus releases more energy when it gains an electron, making the reaction more favorable than B gaining an electron.

c) A will be more to the right on the periodic table than B, since it has a higher ionization energy and larger electron affinity. Because A is further to the right, it has a greater  $Z_{\text{eff}}$  and thus will be smaller than B.

d) Since nonmetals are on the right hand side of the periodic table, and thus have more favorable electron affinity due to a greater  $Z_{\text{eff}}$ , element A is the nonmetal.

## Chemical Bonding Solutions

Problem 71. Answer: B



Formal Charge = Valence electrons – lone pair electrons – bonds

$$FC_F = 7 - 6 - 1 = 0$$

$$FC_P = 5 - 2 - 3 = 0$$

$$FC_S = 6 - 4 - 2 = 0$$

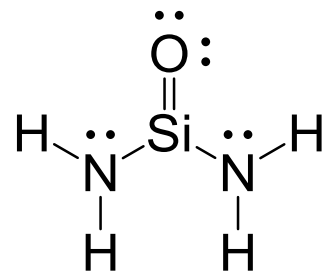
Problem 72. Answer: X is Carbon (C).

$$FC_X = x - 0 - 4 = 0,$$

Therefore  $x = 4$  and that is the number of valence electrons in the element. Giving us carbon.

Problem 73. Answer: C

There are 4 lone pairs of electrons.





## VSEPR Solutions

Problem 74. Answer: C

Solution: A triple bond corresponds to 2  $\pi$  bonds and 1  $\sigma$  bond. The two single bonds consist of one  $\sigma$  bond each. Therefore there is a total of three  $\sigma$  bonds and two  $\pi$  bonds.

Problem 75. Answer:

a) N(1):  $sp^3$    N(2):  $sp^2$    O(1):  $sp^2$    O(2):  $sp^3$

b) i)  $109^\circ$    |   ii)  $120^\circ$    |   iii) 6 lone pairs

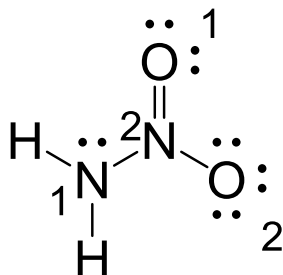
Solution:

N(1) has 3 bonding pairs and 1 lone pair of electrons  $\rightarrow$   $sp^3$  hybridized

N(2) has 3 bonding pairs of electrons  $\rightarrow$   $sp^2$  hybridized

O(1) has 1 bonding pair and 2 lone pairs of electrons  $\rightarrow$   $sp^2$  hybridized

O(2) has 1 bonding pair and 3 lone pairs of electrons  $\rightarrow$   $sp^3$  hybridized



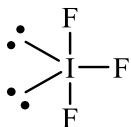
i) N(1) is  $sp^3$  hybridized; tetrahedral structure;  $109^\circ$  bond angle

ii) N(2) is  $sp^2$  hybridized; trigonal planar;  $120^\circ$  bond angle

iii) By the Lewis Structure, there are 6 lone pairs of electrons.

Problem 76. Answer: D

Solution: From the Lewis diagram of the molecule we can see that the shape of the molecule is T-Shaped.

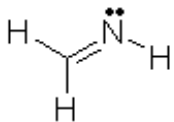


Problem 77. Which of the following statements is/are correct for the formate ion  $\text{HCO}_2^-$  ?

Answer: C

Solution: the oxidation number of the C atom is +2, there are only 2 plausible contributing structures, and  $\text{HCO}_2^- = 18 e^- = \text{AX}_3 = \text{trigonal planar}$

Problem 78. a) Answer:

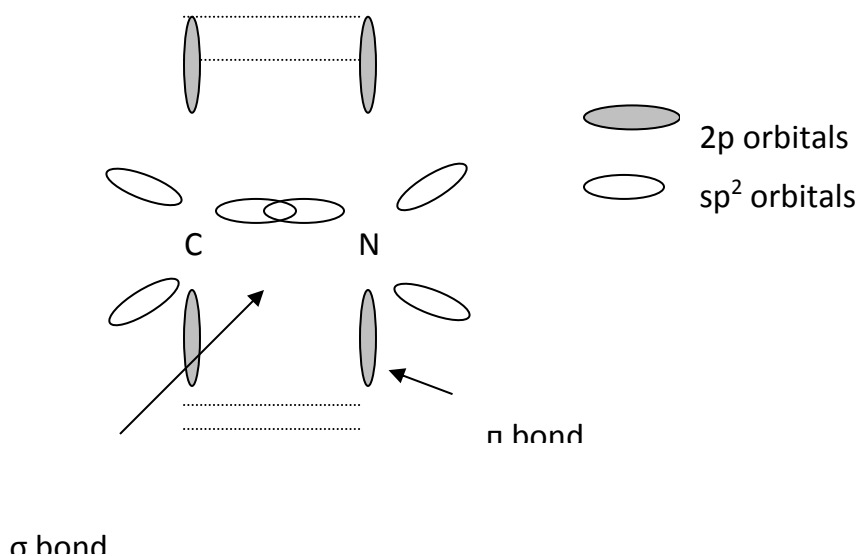


b) Answer:

Sigma bonds = 4

Pi bonds = 1

c) Answer: A single bond contains one  $\sigma$  bond, whereas a double bond consists of one  $\sigma$  and one  $\pi$  bond.  $\sigma$  bonds are covalent bonds between electron pairs in the area between two atoms.  $\pi$  bonds are a covalent bond between parallel p orbitals and the electron pairs shared are below & above the line joining the 2 atoms.

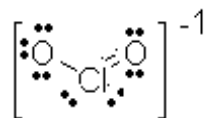


d) Answer: There are 3 electron pairs around the central C atom, and thus it is  $\text{sp}^2$  hybridized. Thus the bond angle is  $120^\circ$ .

Problem 79. Predict the geometric shape of  $\text{ClO}_2^-$  ion.

Answer: Bent

Solution:



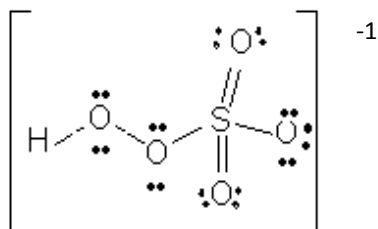
Total valence electrons:  $7 + 2(6) + 1 = 20$  electrons

Center Cl has 2 bonding pairs and 2 lone pairs of electrons  $\rightarrow sp^3$  hybridized & tetrahedral formation. Since there are 2 lone pairs of electrons, the structure is bent-shaped.

Problem 80. Draw the Lewis structure for the peroxymonosulfate ion,  $\text{H-O-O-SO}_3^-$ , and estimate the H-O-O bond angle.

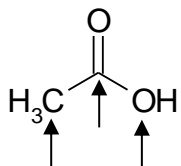
Answer:  $< 109^\circ$

Solution:



The center oxygen in H-O-O has 4 pairs of electrons (2 lone pairs of electron and 2 bonding pairs of electrons); as such, this part of the molecule has a tetrahedral structure. The bond angle is less than  $109.5^\circ$  due to the 2 lone pairs of electrons that bend the structure more than bonding pairs of electrons. The tetrahedral formation means the electrons are distributed in  $sp^3$  orbitals.

Problem 81. a) Applying your knowledge of VSEPR, indicate the geometry of the three atoms indicated in acetic acid.



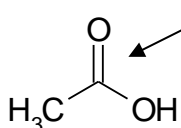
b) What is the  $\text{O}=\text{C}-\text{O}$  bond angle?

c) Indicate the orbital hybridization of the three atoms with the arrows.

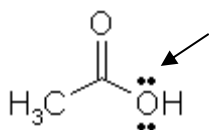
Answer:

a)  $\text{H}_3\text{C}$ - : 4 bonding pairs of electrons, thus it is tetrahedral

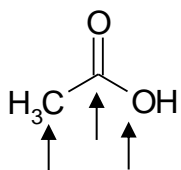
3 bonding pairs of electrons; thus it is trigonal planar



4 bonding pairs of electrons; thus it is tetrahedral.



b) The geometry around the central C carbon is trigonal planar structure. Thus, the bond angle is 120.



c)

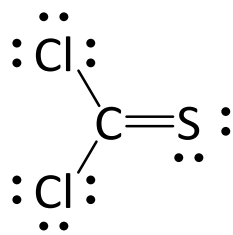
$\text{H}_3\text{C}-\text{C}-\text{OH}$  Tetrahedral structure =  $\text{sp}^3$  hybridization

$\text{H}_3\text{C}-\text{C}-\text{O}$  Trigonal planar structure =  $\text{sp}^2$  planar

$\text{H}_3\text{C}-\text{C}-\text{OH}$  Tetrahedral structure =  $\text{sp}^3$  hybridization

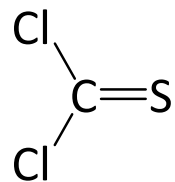
Problem 82. a)  $\text{Cl}_2\text{CS}$  (thiophosgene) (Carbon is central atom, Cl are equivalent).

Lewis:



\*All formal charges are zero. Total valence electrons = 24

VSEPR



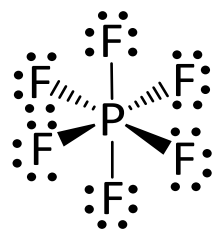
AX3

Electronic geometry: trigonal planar

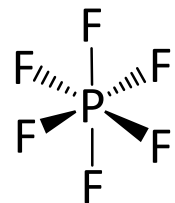
Molecular geometry: trigonal planar

b)  $\text{PF}_6$

Lewis Structure



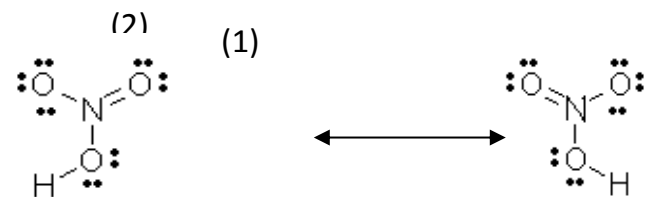
\*Formal Charge on F = 0, Formal Charge on P = -1. Total valence electrons = 48



AX6

Electronic geometry: octahedral

Molecular geometry: octahedral

Problem 83. a)  $\text{HNO}_3$ a) Total # of electrons =  $1 + 5 + 3(6) = 24$ 

Formal charge:

$$\text{On O(1)} \quad 6 - [4 + \frac{1}{2}(4)] = 0$$

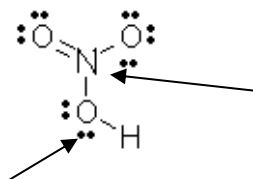
$$\text{On O(2)} \quad 6 - [6 + \frac{1}{2}(2)] = -1$$

$$\text{On O(3)} \quad 6 - [4 + \frac{1}{2}(4)] = 0$$

$$\text{On N} \quad 5 - [0 + \frac{1}{2}(8)] = +1$$

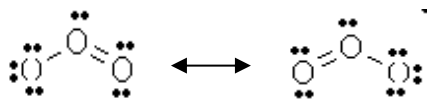
$$\text{On H} \quad 1 - [0 + \frac{1}{2}(2)] = 0$$

Molecular drawing:

Structure around central N: 3  
electron pairs = trigonalStructure around central O: 4  
electron pairs = bentb)  $\text{O}_3$ Total # of valence electrons:  $3(6) = 18$ 

Lewis structure:

(1) (2) (3)

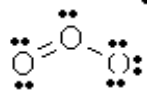


Formal charge on:

$$\text{O(1)} \quad 6 - [6 + \frac{1}{2}(2)] = -1$$

$$\text{O(2)} \quad 6 - [2 + \frac{1}{2}(6)] = +1$$

$$\text{O(3)} \quad 6 - [4 + \frac{1}{2}(4)] = 0$$



Molecular structure around central O = bent