

Chapter 5 Summary – Reactions in Aqueous Solutions

Keywords

Section	Keywords			
<i>Nature of Aqueous Solutions</i>	Nonelectrolyte	Strong electrolyte	Weak electrolyte	
<i>Precipitation Reactions</i>	Precipitate	Net ionic equation	Spectator ions	
<i>Acid-Base Reactions</i>	Acid, Strong acid, Weak acid	Base, Strong base, Weak base	Neutralization	Salt
<i>Oxidation-Reduction</i>	Oxidation, reducing agent	Reduction, oxidizing agent	Redox reaction	Half-reaction
<i>Balancing Redox Reactions</i>	Oxidation number	Disproportionation reactions	Half-reaction method	Total Oxidation State method
<i>Titrations</i>	Titration	Equivalence point	Indicator	

Important Concepts

1. Bronsted-Lowry definition of acid and base: an acid is a proton donator; a base is a proton acceptor. An acid-base reaction is a proton transfer reaction.
2. An oxidation reduction reaction is an electron transfer reaction. An oxidizing agent accepts an electron; a reducing agent donates an electron.
3. Metal oxides react with water to form bases; non-metal oxides react with water to form acids

Tools, Strategies, and Equations

Type of Question	Strategy	Tool/Equation
Electrolyte	If substance is salt, acid, base = electrolyte. Covalent = non-electrolyte. Weak acid, weak base = weak electrolyte	
Precipitation reactions	Determine if a precipitate will form using a solubility chart	Solubility Chart
Acid-Base reactions	Determine if a reaction is acid-base by noting whether there has been a proton transfer	
Redox Reactions	Half-Reaction Method: determine oxidant and reductant from oxidation number changes. Write each half-reaction and balance charge and matter using H ₂ O and H ⁺ or OH ⁻ (whether in acid or base). Add two half-reactions.	
Titrations	Use the equation to determine unknown variable	M ₁ V ₁ = M ₂ V ₂

Chapter 7 Summary – Thermochemistry

Keywords

Section	Keywords				
<i>Terminology</i>	System: open, closed, isolated	Surroundings	Energy, Work	Kinetic energy, Potential energy	Thermal energy
<i>Heat</i>	Heat	Heat capacity, specific heat	Calories, Joules	Law of conservation of energy	
<i>Heats of Reaction and Calorimetry</i>	Chemical energy	Heat of reaction	Exothermic, endothermic	Calorimeter: bomb, coffee cup	
<i>Work</i>	Pressure-volume work				
<i>First Law of Thermodynamics</i>	Internal energy (U)	First Law: $\Delta U = Q + W$	State function	Reversible process, irreversible	
<i>Heats of Reaction</i>	Enthalpy (H)	Standard enthalpy of formation	Standard enthalpy of reaction	Standard state, STP	
<i>Hess's Law</i>	Hess's Law				

Important Concepts

1. Enthalpy is like the total potential energy of a system, and $H = U + PV$
2. Know the ideal gas law, and the concepts of gases: $PV = nRT$
3. Change in enthalpy is equal to the heat added to a system under constant pressure; change in internal energy is equal to the heat added to a system under constant volume
4. If the number of moles of a gas changes from reactants to products, ΔH and ΔU can be related by: $\Delta H = \Delta U + \Delta n_{\text{gas}}RT$ (since $PV = nRT$, $\Delta PV = \Delta nRT$)
5. Enthalpy, internal energy, pressure, volume, temperature, moles are all state functions, i.e. their value depends only on the state of the system. Work and heat are path functions, i.e. their value depends on how the system got to that state.
6. The enthalpy of formation of a pure substance is zero.

Tools, Strategies, and Equations

Type of Question	Strategy	Tools/Equations
Heat released/absorbed	Use the equation to calculate heat in relation to temperature	$Q = C\Delta T = ms\Delta T$
Heat with change in state	Determine how much heat was involved with the change in state; the rest was involved with changing temperature	$Q = n\Delta H^\circ$ $Q = ms\Delta T$

Enthalpy change	Enthalpy equals heat for a constant pressure process; determine heat using relevant equation	$\Delta H = Q_p$ $Q = n\Delta H^\circ$ $Q = ms\Delta T$
Work	Use the equation for work	$W = -P\Delta V$
First Law of Thermodynamics	Use the First Law. If heat or work is not given, use equations to determine	$\Delta U = Q + W$ $W = -P\Delta V$ $Q = ms\Delta T$
Work from chemical equation	If the number of moles of gas increases, the system <i>performed</i> work, i.e. work is negative.	If $\Delta n_{gas} > 0$, $W < 0$ If $\Delta n_{gas} < 0$, $W > 0$ If $\Delta n_{gas} = 0$, $W = 0$
Endothermic/Exothermic from chemical equation	Look for changes in state: solid to liquid, liquid to gas <i>or</i> bond breaking = endothermic. Gas to liquid, liquid to solid <i>or</i> bond forming = exothermic	Also, can calculate ΔH . If $\Delta H > 0$, endothermic If $\Delta H < 0$, exothermic
Determining enthalpy change of reaction from multiple chemical equations	Multiply reactions by proper coefficients, reverse, and combine such that the only remaining molecules form the overall reaction.	Multiplying a reaction by N multiplies the ΔH of that reaction by N. Reversing a reaction multiplies the ΔH by -1
Determining standard reaction enthalpy from chemical equation and standard enthalpies of formation	Sum the enthalpies of formation of each product times its stoichiometric coefficient. Subtract the enthalpies of formation of each reactant times its stoichiometric coefficient.	Standard Enthalpy of Formation Table, $\Delta H^\circ = \sum n[p] * \Delta H_f(p) - \sum n[r] * \Delta H_f(r)$

Chapter 8 & 9 Summary – Quantum Theory and Periodic Trends

Keywords

Section	Keywords			
<i>Electromagnetic Radiation</i>	Wavelength	Frequency	Amplitude	Speed of light
<i>Quantum Theory</i>	Quantum	Planck's constant	Photoelectric effect	Photons
<i>The Bohr Atom</i>	Quantum number, n	Ground state	Excited state	Energy-level diagram
<i>Quantum Numbers & Electron Orbitals</i>	Principal quantum number, n (size), "shell"	Orbital angular momentum, l (shape), "subshell"	Magnetic quantum number, m_l (orientation)	Electron spin number, m_s
<i>Electron Configuration</i>	Pauli Exclusion Principle	Hund's Rule	Aufbau Principle	spdf notation, orbital diagram
<i>Sizes of Atoms and Ions</i>	Atomic radii	Effective nuclear charge	Ionic radii	Isoelectronic
<i>Ionization Energy & Electron Affinity</i>	Ionization energy, IE	Electron affinity, EA		
<i>Magnetic Properties</i>	Paramagnetic	Diamagnetic		
<i>Periodic Properties of Elements</i>	Metals, non-metals, metalloids	Acid anhydride (non-metal oxide)	Base anhydride (metal oxide)	Amphoteric behaviour

Important Concepts

1. 4 quantum numbers specify a unique electron in an atom. By the Pauli Exclusion Principle, no two electrons can share the same four quantum numbers.
2. Energy is quantized in electron orbitals and light. Electrons can jump from the ground state to an excited state by the absorption of light, or vice versa by the emission of light.
3. Be sure to consider the Pauli Exclusion Principle, Hund's Rule, and the Aufbau principle when writing electron configurations.
4. Atomic radii, electron affinity, and ionization energy are all based on the effective nuclear charge, and the effects of filled and half-filled subshells (special stability).

Tools, Strategies, and Equations

Type of Question	Strategy	Tools/Equations
Wave question	Use the wave equation	$c = \lambda v$
Energy of photon	Use energy of photon equation	$E = hv = hc/\lambda$
Photoelectric effect	Use the equation for photoelectric effect	$E_{\text{photon}} = \phi + \frac{1}{2}mv^2$
Calculate the energy of an electron in an orbital	Use the Rydberg formula	$En = -R\left(\frac{Z^2}{n^2}\right)$

		Where Z is the effective nuclear charge, and R = -2.18 E -18
Calculate an electron emission or absorption	Use the energy of the photon equal to the difference in orbital electron energy	$E_{\text{photon}} = \Delta E_n$
Atomic Radii ordering	Use fact that atomic radii decreases with effective nuclear charge.	General Trend: increasing down and to the left
Ionization Energy ordering	Ionization energy increases with effective nuclear charge. Drops slightly after a half-filled orbital, drops largely after a filled	General trend: increasing up and to right
Electron affinity	Increases with increasing effective nuclear charge. Low in magnitude at half-filled and filled subshells.	General trend: increasing in magnitude up and to right.

Chapter 10 Summary – Lewis Structures and Bonding

Keywords

Section	Keywords			
<i>Lewis Theory</i>	Lewis structure	Octet rule	Ionic bonds	Covalent bonds
<i>Covalent Bonding</i>	Bond pairs, lone pairs	Single, double, triple	Coordinate covalent	
<i>Polar Covalent Bonds</i>	Polar covalent bond	Electronegativity (EN)	Electronegativity difference	
<i>Writing Lewis Structures</i>	Skeletal structure	Central atom	Terminal atom	Formal charge
<i>Resonance</i>	Resonance			
<i>Exceptions to Octet Rule</i>	Free radicals (paramagnetic)	Incomplete octets (Group II, III)	Expanded valence shells (>3 rd period)	
<i>Bond Order and Bond Lengths</i>	Bond order (single, double, triple)	Bond length		
<i>Bond Energy</i>	Bond dissociation energy	Average bond energies	Lattice energy	
<i>Shapes</i>	VSEPR	Electron geometry vs molecular geometry	Bond angle	Dipole moment and polar molecules

Important Concepts

1. Bonds are not either ionic or covalent; rather, they have more or less ionic/covalent tendencies. A large difference in EN between atoms in a bond = more ionic; A small difference in EN = more covalent; a middling difference in EN = polar covalent.
2. Electronegativity is a measure of how strongly an atom pulls electrons in a bond. $EN = IE - EA$
3. When writing Lewis structures, if there are no elements beyond the 2nd period, the octet rule trumps formal charge. If there is an element in 3rd period or beyond, expanded valence shells can be used to minimize formal charge.
4. Bond energies can be used to predict the ΔH of reaction, but they are less accurate than using enthalpies of formation, as they are averages over different molecules.
5. Lattice energy depends on both the charge of the ions and their atomic radii.
6. The shape of molecules and their polarity can be described by valence shell electron repulsion theory, which states that electron domains will be as far from each other as possible.

Tools, Strategies, and Equations

Type of Question	Strategy	Tool/Equation
Evaluating validity of a Lewis Structure	Check to see if octet rule and formal charge rules have been	Octet Rule Formal Charge Rules:

	violated.	<ol style="list-style-type: none"> 1. Sum of FC = charge 2. Minimize FC 3. Negative FC on most EN 4. Same sign FC on adjacent = unlikely
Calculating Formal Charge	Use formal charge rules. A formal charge is assigned to each atom	$FC = \text{Group } \# - \# \text{ of electrons} - \# \text{ of bonds}$
Drawing Lewis Structures	<ol style="list-style-type: none"> 1. Place the least EN atom at centre 2. Draw the skeletal structure 3. Place multiple bonds such that FC is minimized, and FCs add to charge on molecule 4. If a 3rd period element is present, an expanded valence may be used 5. Draw all possible resonance structures 	Electronegativity principles Formal charge Octet Rule Expanded Valence Resonance
Average Bond Order	In a resonance hybrid, calculate average bond order using one resonance structure. Add all bond orders of bonds which are switched between resonance structures, and divide by number of atom pairings in same.	Resonance Bond Order
Average formal charge	Calculate the average formal charge of atoms in a resonance hybrid in a similar way to average bond order. Add all formal charges, divide by number of bond-changed species	Resonance Formal charge
Enthalpy of reaction (covalent molecules) with bond energy	<ol style="list-style-type: none"> 1. Lewis structure of reactants and products 2. Determine which bonds have been broken and formed 3. Use Hess's law, along with bond dissociation energy, to determine ΔH of reaction 	Lewis Structures Bond dissociation energy Hess's Law for BE: $\Delta H = \sum BE[rb] - \sum BE[pb]$

	4. Incorporate the ΔH of any phase transitions (reaction must occur in gas phase)	
Born-Haber Cycle enthalpy (ionic molecules)	1. Write out formation equation 2. ΔH for phase transitions/bond breaking 3. IE of cation 4. EA of anion 5. ΔH of lattice for ionic structure	Bond breaking Lattice energy IE EA
Determine shape of molecule	1. Draw Lewis structure 2. Determine electron geometry from AX_nE_m notation 3. Determine molecular geometry from same 4. Determine if bond angles are ideal	AX_nE_m notation Electron geometries Molecular geometries
Determine if molecule is polar	1. Draw Lewis structure 2. Draw dipole moments for each bond 3. Determine if all dipole moments cancel or add up to a molecular dipole moment	ΔEN Dipole concept $\mu = \delta r$

Chapter 15 Summary – Equilibrium

Keywords

Section	Keywords			
<i>Dynamic Equilibrium</i>	Dynamic equilibrium			
<i>The Equilibrium Constant</i>	Equilibrium constant (K)	Activity		
<i>The Reaction Quotient Q</i>	Reaction quotient (Q)	Direction of equilibrium shift		
<i>Le Chatelier's Principle</i>	Le Chatelier's Principle			
<i>Equilibrium Calculations</i>	ICE table			

Important Concepts

1. The equilibrium constant varies only with temperature; no other change in an equilibrium system will change the equilibrium constant.
2. Large (10^{10}) K = reaction goes to “completion”; small (10^{-10}) K = reaction does not occur significantly; intermediate K implies an equilibrium.
3. The reaction quotient can be used to predict the value of an equilibrium shift by comparing it to the equilibrium constant
4. Le Chatlier’s Principle states that if a change is made to a system in a state of dynamic equilibrium, the system will try to shift to a new position of equilibrium to minimize the effect of the change.
5. Catalysts have no effect on the position of equilibrium

Tools, Strategies, and Equations

Type of Question	Strategy	Tools/Equations
Writing an equilibrium constant	Use the definition of the equilibrium constant (keep in mind to only include aqueous solutions and gases, not solids or liquids)	$K = \frac{a_{p1}^{np1} * a_{p2}^{np2} * \dots}{a_{r1}^{nr1} * a_{r2}^{nr2} * \dots}$
Solving for a single unknown in a state of equilibrium	Use the definition of K and solve for the unknown	Definition of K (above)
Add reactions and find the net equilibrium constant	- K reverse is the reciprocal of K forward - If a reaction is multiplied by a constant, K is raised to the power of that constant	$K_{reverse} = \frac{1}{K_{forward}}$ $K_{net} = K^n$

	<p>- When two reactions are coupled, the K for the overall reaction is the product of their individual Ks</p>	$K_{net} = K_1 * K_2 * \dots$
Predicting the direction of equilibrium change from given activities	Compare the reaction quotient to the equilibrium constant	If $Q < K$, reaction proceeds forward If $Q > K$, reaction proceeds in reverse If $Q = K$, reaction is in equilibrium
Predicting the equilibrium activities (concentration or pressure) of species in a reaction given initial activities and equilibrium constant	Use an ICE table to obtain an expression for K in terms of known values and the unknown value, and set equal to the given K. Solve for the unknown.	ICE table
Predicting the direction of equilibrium shift given a change in the system	Use Le Chatelier's Principle to predict the shift in equilibrium. Note that a catalyst will not shift the position of equilibrium. Also, only temperature will change the equilibrium constant.	

Chapter 19 Summary – Entropy and Gibbs Free Energy

Keywords

Section	Keywords			
<i>Spontaneity</i>	Spontaneous process	Nonspontaneous process		
<i>The Concept of Entropy</i>	Entropy	Entropy change in terms of heat		
<i>Evaluating Entropy and Entropy Changes</i>	Third law of thermodynamics	Standard Molar Entropy	Trouton's rule	
<i>Criteria for Spontaneous Change</i>	Second law of thermodynamics	Gibbs Free Energy (G)	Gibbs Free Energy Change (ΔG)	
<i>Standard Free Energy Change</i>	Standard free energy change	Standard molar free energies of formation		
<i>Free Energy Change and Equilibrium</i>	Thermodynamic equilibrium constant			
<i>Coupled Reaction</i>	Coupled reactions			

Important Concepts

1. Entropy is measure absolutely, as per the third law of thermodynamics
2. The Gibbs free energy of any spontaneous process is less than zero.
3. The equilibrium constant for a particular chemical reaction can be related to the change in Gibbs free energy of that reaction and the temperature at which the reaction occurs.
4. A nonspontaneous reaction can occur if coupled with a spontaneous reaction such that the sum of their respective change in Gibbs free energy is less than zero.
5. The change in Gibbs free energy is defined by: $\Delta G = \Delta H - T\Delta S$. Below the crossover temperature T_{cross} , ΔH controls the reaction; above T_{cross} , ΔS controls reaction. The crossover temperature is given by $T_{cross} = \frac{\Delta H}{\Delta S}$

Tools, Strategies, and Equations

Type of Question	Strategy	Tool/Equation
Determine whether a process increases or decreases entropy	Look for changes in the number of moles of gas, liquid, or aqueous solution.	If $\Delta n_{gas,liquid,aq} > 0, S > 0$
Determine the change in entropy for a transfer of heat	The change in entropy of the system is equal to the heat transferred divided by the system's temperature	$\Delta S = \frac{q_{sys}}{T}$

Determine the standard free energy change given standard free energies (same for standard entropy change)	Use the formula, essentially identical to the change in enthalpy equation.	$\Delta G^\circ = \sum n[p]\Delta G^\circ[p] - \sum n[r]\Delta G^\circ[r]$
Determine the standard free energy change (or entropy) given a series of reactions	Add changes in free energy as per Hess's Law	
Determine the equilibrium constant for a reaction given the change in free energy (or vice-versa)	Use the formula to solve for the appropriate variable	$\Delta G^\circ = -RT\ln(K)$
Determine whether a reaction is spontaneous	Calculate the change in free energy, and see if it is less than zero.	$\Delta G = \Delta H - T\Delta S$
Determine the change in free energy at non-standard conditions	Use the formula, where Q is the reaction quotient.	$\Delta G = \Delta G^\circ + RT\ln(Q)$ $\Delta G = RT\ln(\frac{Q}{K})$
Determine if a coupled reaction will work.	Determine the change in free energy for each, and add the ΔGs. If the sum is less than zero, the coupled reaction will work.	

Chapter 20 Summary - Electrochemistry

Keywords

Section	Keywords			
<i>Electrode Potentials</i>	Electrochemical cell	Electrodes: Anode, Cathode	Cell diagram	Cell potential (E_{cell})
<i>Standard Electrode Potentials</i>	Standard hydrogen potential	Standard cell potential (E°_{cell})		
E_{cell} , ΔG , and K	Faraday constant (F)	Relationship between E_{cell} and ΔG		
E_{cell} as a function of concentration	Nernst equation	Spontaneous, nonspontaneous	Concentration cell	
Batteries	Battery	Primary cells	Secondary cells	Fuel cells
Corrosion	Cathodic protection			

Important Concepts

1. Electrochemical cells are created as a result of some metals being better oxidizing agents than others.
2. A cell diagram shows the anode, cathode, and salt bridge.
3. The standard cell potential can be obtained by adding the half-cell potentials. That is, $E^{\circ}_{cell} = E_{\text{reduction}} + E_{\text{oxidation}}$
4. E_{cell} can be determined for nonstandard conditions using the Nernst equation, which relates E to the reaction quotient Q.
5. Cathodic protection involves attaching a metal which is a better reducing agent than another metal to that same metal in order to protect it; for example, zinc is attached to a piece of iron because zinc is a stronger reducing agent than iron.

Tools, Strategies, and Equations

Type of Question	Strategy	Tool/Equation
Determine the E°_{cell} of an electrochemical cell given its half-cell reactions	Add the standard cell potentials of the half-cell reactions, given that for a particular half-reaction, $E_{\text{oxidation}} = -E_{\text{reduction}}$	$E^{\circ}_{cell} = E^{\circ}_{\text{red.}} + E^{\circ}_{\text{ox.}}$
Predict whether an oxidation reaction will occur	Determine the E°_{cell} of the reaction. If $E > 0$, reaction is spontaneous. If $E < 0$, reaction is nonspontaneous.	
Write the cell reaction given the cell diagram	Know the convention for cell diagram notation: left side is anode, right is cathode	$M_{(s)} M^{x+}_{(aq)} N^{y+}_{(aq)} N_{(s)}$

Determine the ΔG° given the E_{cell}°	Use the equation and solve.	$\Delta G = -nFE_{cell}^\circ$ F is Faraday's constant, 96 485 C/mol e ⁻
Determine the E_{cell} of an electrochemical cell under nonstandard conditions	Use the equation and solve.	$E_{cell} = E_{cell}^\circ - \frac{RT}{nF} \ln(Q)$ Or $E_{cell} = E_{cell}^\circ - \frac{0.0257}{n} \ln(Q)$ At 298 K
Determine the best metal to use for Cathodic protection	Say you want to protect metal "A". Select a metal which is a stronger reducing agent (oxidized more easily) than "A."	Concept of Cathodic protection

Chapter 16 Summary – Acid/Base

Keywords

Section	Keywords			
<i>Arrhenius Theory</i>	Arrhenius acid	Arrhenius base		
<i>Bronsted-Lowry Theory of Acids and Bases</i>	Acid = proton donor	Base = proton acceptor	Conjugate acid/base	Acid/Base ionization constant (K_A & K_B)
<i>Self-Ionization of Water and the pH Scale</i>	Self-ionization of water	Water equilibrium constant (K_w)	pH and pOH	
<i>Strong Acids and Strong Bases</i>	Ionization of strong acids	Ionization of strong bases		
<i>Weak Acids and Weak Bases</i>	Acid/Base ionization constant (K_A & K_B)	pK_A and pK_B	Percent ionization	
<i>Polyprotic Acids</i>	Polyprotic acids			
<i>Ions as Acids and Bases</i>	Hydrolysis reactions	Salts as weak acids/bases		
<i>Molecular Structure and Acid Base Behaviour</i>	Bond strength of H^+ ions in acids	Lone-pair availability in bases		
<i>Lewis Acids and Bases</i>	Lewis Acid	Lewis Base	Adduct	

Important Concepts

1. There are three definitions of acids and bases:
 - a. Arrhenius: Acids create H^+ ions in solution; Bases create OH^- .
 - b. Bronsted-Lowry: Acids are proton donators; Bases are proton acceptors.
 - c. Lewis: Acids accept lone-pairs; Bases donate lone-pairs.
2. Water will self-ionize, thus the concentration of H^+ and OH^- in water will always be related to the water's equilibrium constant, $K_w = 1.0 \times 10^{-14}$. This leads to some important relationships:
 - a. $K_w = [H^+][OH^-]$
 - b. $K_w = K_A K_B$
 - c. $pK_w = pK_A + pK_B$
3. Strong acids and bases dissociate completely in water; weak acids and bases dissociate based on their equilibrium constants, K_A and K_B respectively.
4. Acids which contain multiple hydrogen ions are referred as polyprotic acids because they can ionize multiple times.
5. The weaker the bond strength of an H^+ ion in a molecule, the stronger the acid; the more weakly a lone-pair is held in a base, the stronger the base.
6. The addition compound of a Lewis acid-base reaction is called an adduct.

Tools, Strategies, and Equations

Type of Question	Strategy	Tool/Equation
Determine the conjugate acid and base from a chemical equation	Determine which substance has accepted a proton and which has donated one.	Concept of Bronsted-Lowry acids/bases
Calculate the pH, pOH concentration of H ⁺ and/or OH ⁻ , etc. of a strong acid or base dissolved in water	Note that a strong acid or base will dissociate completely in water, thus the ion concentration is easily determined. The other ion can be determined from K _w relation	$K_w = [H^+][OH^-]$
Calculate the pH, pOH concentration of H ⁺ and/or OH ⁻ , etc. of a weak acid or base dissolved in water	Use an ICE table, along with the given K _A or K _B to determine ion concentration	$K_w = K_A K_B$ $K_A = \frac{[H^+][A^-]}{[HA]}$
Calculate the percent ionization of a weak acid/base	Use the above strategy, along with the equation for percent ionization	$PI = \frac{[H^+]_{final}}{[HA]_{initial}} \times 100\%$
Calculate the pH, pOH concentration of H ⁺ and/or OH ⁻ , etc. of a salt dissolved in water	If the cation is a weak acid, or the anion is a weak base, the dissolved salt will act as an acid or base. Calculate using the above strategies	
Determine which among two acids or bases is stronger based on molecular formula	Draw the Lewis structure. For acids, the weaker the H ⁺ bond, the stronger the acid. Weakness is caused by highly electronegative atoms pulling electrons away from the hydrogen (O, Cl, Br, I, etc.). For bases, the more available the lone pair, the stronger the base. Availability is cause by a lack of highly electronegative atoms pulling the electrons away (O, Cl, Br, I, etc.).	Concept of EN Lewis structures
Determine, based on a chemical equation, which molecule is the Lewis acid/base.	Draw Lewis structures for reactants and products. The Lewis acid is the molecule missing a lone pair; the Lewis base is the molecule with a free lone pair.	Concept of Lewis acids/bases Lewis structures