Review: Energy Changes

1. Define Enthalpy
2. List the three categories of enthalpy changes and give examples
3. For each of the following thermochemical equations…
	* Rewrite the equation with a ΔH value
	* State whether the reaction is endothermic or exothermic
	* Draw a potential energy diagram for the process
	1. Ag(s) + ½ Cl2(g) → AgCl(s) + 127.0 kJ
	2. H2O(l) + 44.0 kJ → H2O(g)
4. What is “Q?”
5. Define specific heat capacity. What are the units for specific heat capacity?
6. State the first and second laws of thermodynamics.
7. Explain the difference between the terms system and surroundings.
8. Identify the system and surroundings in the following situations
	1. 50 mL of 1.0 mol/L NaOH is used to neutralize 50 mL of 1.0 mol/L HCl
	2. Sugar is dissolved in hot tea
9. What assumptions do we make when calculating the enthalpy of a reaction that occurs in solution?
10. Explain the difference between an open, closed and isolated system.
11. Explain why you need to change the sign of ΔH when you reverse an equation.
12. Explain why it is important to include the state when writing thermochemical equations.

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Review: Energy Changes Answers

1. Define Enthalpy – the total energy of the system plus the pressure times the volume H = E + PV
2. List the three categories of enthalpy changes and give examples

Enthalpy of solutions NaOH(s) 🡪 Na+(aq) + OH-(aq)

Enthalpy of phase changes H2O(g) 🡪 H2O(l)

Enthalpy of chemical reactions CH4(g) + O2(g) 🡪 CO2(g) + 2H2O(g)

1. For each of the following thermochemical equations…
	* Rewrite the equation with a ΔH value
	* State whether the reaction is endothermic or exothermic
	* Draw a potential energy diagram for the process
	1. Ag(s) + ½ Cl2(g) → AgCl(s) + 127.0 kJ

 Ag(s) + ½ Cl2(g) → AgCl(s ΔH = - 127.0 kJ exothermic

* 1. H2O(l) + 44.0 kJ → H2O(g)

H2O(l) → H2O(g) ΔH = 44.0 kJ endothermic

1. What is “Q” – the amount of heat (J) Q = mcΔT
2. Define specific heat capacity. What are the units for specific heat capacity? Specific heat capacity is the amount of energy required to raise the temperature of 1g of a substance by 1ºC.
3. State the first and second laws of thermodynamics.

First Law of Thermodynamics – energy can be converted from one form to another but cannot be created or destroyed Esystem = - Esurroundings

Second Law of Thermodynamics – when two objects are in thermal contact, heat is always transferred from the object at a higher temperature to an object at a lower temperature until the two objects are the same temperature

1. Explain the difference between the terms system and surroundings.

System – reaction, dissolving etc.

Surroundings – everything around it

1. Identify the system and surroundings in the following situations
	1. 50 mL of 1.0 mol/L NaOH is used to neutralize 50 mL of 1.0 mol/L HCl system – reaction of NaOH and HCl, surroundings water, air, beaker etc.
	2. Sugar is dissolved in hot tea system – tea and sugar, surroundings – air, water, mug
2. What assumptions do we make when calculating the enthalpy of a reaction that occurs in solution?
The system is isolated – no heat is transferred between the calorimeter and the outside environment

Any heat absorbed or released by the calorimeter materials, such as the container, is negligible

A dilute solution is assumed to have the same volume, density, and specific heat capacity as pure water

The process takes place under constant pressure

1. Explain the difference between an open, closed and isolated system.

Open system – can exchange both energy and matter with its surroundings

Closed system – can exchange energy, but not matter with its surroundings

Isolated system – cannot exchange energy or matter with its surroundings

1. Explain why you need to change the sign of ΔH when you reverse an equation. If the forward reaction is exothermic, the reverse reaction is endothermic
2. Explain why it is important to include the state when writing thermochemical equations. Each state has different energy values.

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